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TEXT-BOOK

ELEMENTARY CHEMISTRY,

FOR THE USE OF

SCHOOLS AND JUNIOR STUDENTS,

(WITH NEW NOTATION, ETC.)

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CHEMISTRY.

CHAPTER I.

INTRODUCTION

1. CHEMISTRY teaches us the nature and properties of the different substances of which the earth, the air, and the sea, as well as all animal and vegetable bodies, are composed.

2. It also shows us the changes which these substances are continually undergoing; the manner in which such changes are produced; and the laws by which they are regulated.

3. All matter, by which is meant everything that occupies space, is composed of very minute particles or atoms, united together by an attractive force, termed attraction of cohesion.

4. Attraction of cohesion is the power whereby particles of the *same* kind of matter are held together and united into a mass having the *same* properties as the particles of which it is formed.

5. Thus, two drops either of water, oil, or quicksilver, when brought into contact, will run together into one mass; but, however much these three substances may be shaken together, they will soon separate when left at rest.

It is through this attraction of cohesion that a drop of water is always spherical; and that small particles of mercury are constantly of a globular figure. If a plate of glass be laid upon a globule of mercury, and loaded with weights, one after another, the mercury is pressed out into a thin layer; but as soon as the weights are removed, its globular figure is restored.

6. Matter exists in three different states, viz., *solid*, *liquid*, or *aëriform*. These different states, or, as they have been termed, physical conditions of matter, depend on the different degrees of cohesive force exerted between its particles.

7. The force of cohesion in a substance increases in proportion as its particles are brought nearer together.

8. The more firmly the particles of a body cohere, the harder it is.

Hence, whatever brings the particles of a body closer together, tends to harden it.

9. Common clay is soft, plastic, and contains much moisture; but when strongly heated, most of this moisture is driven out, and the particles of the clay are brought into closer contact. The mass thus becomes hard, as we find it in bricks, tiles, and earthenware.

10. When the particles of which a body is composed cohere firmly, so as to resist impression, the body is said to be in a *solid* state.

"The cohesive attraction which exists between the particles of a solid body may be measured by the weight necessary to overcome it. Thus, if a rod of metal, glass, wood, &c., be suspended in a perpendicular direction, and weights attached to its lower extremity till the rod break, the whole weight thus attached is the measure of the cohesive force, or *tenacity* of the rod."

11. When the particles of which a body is composed do not cohere firmly, but move freely in any direction among

themselves, and hence yield to the slightest impression, the body is said to be in a *fluid* state.

12. Fluids may be divided into *elastic* and *non-elastic* fluids. They are also spoken of as *compressible* and *incompressible.*

13. Air, and all *aëriform* bodies, whether gases or vapors, belong to the former class of elastic, compressible fluids.

Note.—Aëriform means *in the form of air.*

14. Water and all other *liquids* belong to the class of incompressible, non-elastic fluids.

15. Strictly speaking, no known fluid is *wholly* incompressible and non-elastic, but all ordinary liquids are sufficiently nearly so, for us to regard them as such, without any sensible error.

16. Matter exists in the *liquid* state when its particles cohere slightly, but not with sufficient force to prevent them separating by the influence of their weight alone.

17. Matter exists in an *aëriform* state when the particles composing it have a mutual repulsion, in consequence of which it increases in volume or bulk unless restrained by some external pressure.

18. Heat is the great antagonist to cohesion. As the latter draws the particles of matter together, so the former tends to drive them asunder. Hence the increase or diminution of heat is the chief cause of the different states in which matter exists.

19. When the cohesion between the particles of a body is strong, the body is a solid;* when the cohesion is neu-

* Lavoisier thus explained solidity. "The particles of all bodies may be considered as subject to the action of two opposite powers *attraction* and *repulsion*. So long as the attractive force remains the stronger, the body must remain in a state of solidity; but if, on the contrary, heat has so far removed these particles from each other as to place them beyond the sphere of attraction, they lose the cohesion they before had, and the body ceases to exist in the solid state."

tralised, or, at least, very considerably weakened, by the influence of an increased temperature, the body becomes liquid; and when by a still greater increase of heat the cohesion is altogether destroyed, and the particles begin to repel each other, the body assumes the state of gas or vapor.

20. A *gas* is a permanently elastic fluid, not changeable into a liquid by the *ordinary* pressure and temperature of the atmosphere, as the air.

21. A *vapor* is an elastic fluid which is changeable into a liquid by ordinary cold, as steam.

22. The operation of changing the body from the aëriform to the liquid state is termed *condensation.*

23. It is to be observed that although some gases are permanently elastic under all the circumstances in which they have been experimented upon, yet the application of unusual pressure or extreme cold to other gases,* changes them into the liquid and even into the solid state.

24. Many substances are capable of appearing under all three forms of matter.

Of this water affords a familiar illustration; for it is solid, in the form of ice; liquid, as water; and in an aëriform state, when converted into steam, or vapor.

25. Mercury presents another example; it is liquid at ordinary temperatures, but if cooled to 39° below 0° of Fahrenheit's thermometer, it becomes solid; on the other hand, if it be heated to 662° it *volatilizes,* or becomes converted into vapor.

26. Other bodies cannot be made to present such changes, and hence appear only under two forms. Iron,

* If the strict distinction between a gas and a vapor, namely, that a gas cannot be condensed while a vapor can, is to be rigidly observed, these bodies, usually termed gases, must be termed vapors, and their number will probably be increased as more powerful modes of condensation are discovered.

for example, though capable of being melted, and thus taking the liquid form, has never been exhibited in the state of vapor. Ether, on the other hand, though a liquid which readily evaporates, has not yet been solidified or frozen.

27. It is, however, believed, that it would be possible to cause all substances to assume the three physical conditions of matter, by the action of greater heat or cold than can be at present produced, provided such substances are not destroyed,* or, more strictly speaking, decomposed by the increase or diminution of temperature.

28. All bodies, whether solid, liquid, or aëriform, are divided by the chemist into two classes, *simple* and *compound.*

29. Simple bodies are such as cannot be separated into parts having different properties.

30. Compound bodies being made up of simple bodies, can be *decomposed* or divided into their elementary or component parts.

31. "Thus brass is a compound body, and may be shown to consist of copper and zinc; but neither copper nor zinc can be separated into any other substances: they may be ground, crushed, melted, dissolved, over and over again, but the copper can only be made to yield copper, and the zinc, zinc." These two metals are therefore termed simple bodies or elements.

32. It is not meant to be said that all the bodies now called elements are really so, but only, that up to the present time, they have resisted every effort to decompose them.

* "It is a fundamental principle of all physical knowledge, that we can neither create anything, nor destroy anything. We may change the fashion and properties of things; but to form new laws of combination, or new species of matter, belongs to the Creator alone. —*Francis.*

33. Fire, air, earth, and water, were for ages regarded as elements: but modern science has proved the latter three to be compounds, and the first to be an effect of intense chemical action.

34. The decomposition or *taking to pieces* of a compound body, in order to discover its constituent elements, is called *Analysis*, from two Greek words, *ana*, thoroughly, and *lusis*, a loosening.

35. The formation of a compound by putting together the elements of which it is composed, is termed *Synthesis*, also from two Greek words, *sūn*, together, and *thesis*, a placing.

36. The elements are not all of equal importance, for some are found only in very minute quantities in a few scarce minerals, while others are continually occurring in various combinations, to form the atmosphere, and the ocean, together with the great rocky masses of the earth's crust, and the organized bodies which live upon its surface.

37. By *Organized* bodies, we mean animals and plants, which live and grow by the help of what are called *Organs*, or leaves, roots, lungs, heart, stomach, &c.

38. The products formed by these Organs, such as sap, sugar, starch, gum, blood, &c., are termed *organic* substances; and hence the chemistry of plants and animals is called *Organic Chemistry*.

39. Air, water, and minerals, on the other hand, are neither produced, nor continued in being, by means of organs; they are therefore called *inorganic* bodies, and the branch ot chemistry relating to them is termed *Inorganic Chemistry*.

EXERCISES ON CHAPTER I.

1. What does Chemistry teach us?
2. Of what changes, &c., does it treat?
3. Of what is matter composed?
4. What is attraction of cohesion?

5. Give illustrations of the attraction of cohesion.
6. Name the different states of matter, and on what they depend.
7. When does the cohesive force between the particles of a body increase?
8. When is one body harder than another?
9. Shew how clay is hardened by burning.
10. When is a body in a *solid* state?
11. When is a body in a *fluid* state?
12. How may fluids be arranged?
13. What bodies belong to the class of *elastic* fluids?
14. To which class do all *liquids* belong?
15. Why is this arrangement of fluids not *strictly* correct?
16. When is a body in a liquid state?
17. When does matter exist in an aëriform state?
18. How is heat the antagonist of cohesion?
19. How are the physical conditions of matter influenced by temperature?
20. What is a gas?
21. Define a vapor.
22. What is meant by condensation?
23. What are the chief agents in condensing aëriform bodies?

24. Many substances are capable of appearing under all three conditions of matter. Give a familiar example.

25. How does mercury illustrate the same fact?

26. Give examples of bodies capable of appearing under only two forms.

27. What is believed in reference to bodies which appear under only two forms?

28. How are all bodies arranged by the chemist?

29. What is a Simple body or Element?

30. What is a Compound body?

31. Give an example of a simple and of a compound body

32. With what limitation is the word Element to be used?

33. Why do we no longer regard the ancient division of matter into four elements?

34. What is Analysis?

35. What is Synthesis?

36. Why are some of the elements more important than others?

37. What is meant by Organized bodies?

38. What is Organic Chemistry? Name some organic products

39. What is Inorganic Chemistry? Why are some bodies term. *inorganic* bodies? Give examples.

CHAPTER II.

CHEMICAL NOMENCLATURE, NOTATION, AND SYMBOLS.

1. THE elementary substances at present known are sixty-five in number, and may be conveniently arranged in two groups.

2. The first group comprises thirteen substances, which have been termed Non-metallic elements or Metalloids.

3. The second group includes the remaining fifty-two bodies which are known as Metals. Several of these are of very recent discovery, and their properties are as yet very imperfectly understood.

4. The following is an alphabetical list of the elements in each group, with the letters or symbols by which, for the sake of convenience, they are usually distinguished.*

5. The symbol of each element consists of the first letter of its name, in Capital. When the names of several elements begin alike, the single capital is reserved for the first discovered, or more important element; the others being designated by their initial letter conjoined with a small one, usually the most characteristic in the word.†

6. The names of the elements (with the exception of the common metals and one or two others) are generally expressive of some leading property. Thus, chlorine ‡ takes its name from its greenish colour, and iodine from its violet vapor. All the recently-discovered metals end in *um*.

* The most important have been printed in black type, and those of rare occurrence in *italics*.

† Sometimes the symbol is derived from a Latin name, having an initial different from that of its ordinary name. In such cases, the Latin name has been given in parentheses.

‡ Chlorine, from χλωρος, greenish yellow; iodine, from ιον, a violet.

TABLE OF ELEMENTARY SUBSTANCES.

EQUIVALENT POWER.	NAMES OF THE ELEMENTS.	SYMBOLS.	COMBINING NUMBERS.
	METALLOIDS.		
Monatomic ..	Bromine	Br.	80
,, ..	Chlorine	Cl.	35.5
,, ..	Fluorine	F.	19
,, ..	Hydrogen	H.	1
,, ..	Iodine	I.	127
Diatomic ..	Oxygen	O.	16
,, ..	*Selenium*	Se.	79.5
,, ..	Sulphur	S.	32
Triatomic ..	Boron	B.	11
,, ..	Nitrogen	N.	14
,, ..	Phosphorus	P.	31
Tetratomic ..	Carbon	C.	12
,, ..	Silicon	Si.	28
	METALS.		
Monatomic ..	*Cœsium*	Cs.	133
,, ..	*Lithium*	Li.	7
,, ..	Potassium (Kalium) ..	K.	39.1
,, ..	*Rubidium*	Rb.	85.4
,, ..	Silver (Argentum) ..	Ag.	108
,, ..	Sodium (Natrium) ..	Na.	23
,, ..	*Thallium*	Tl.	204
Diatomic ..	Barium	Ba.	137
,, ..	Cadmium	Cd.	112
,, ..	Calcium	Ca.	40
,, ..	*Cerium*	Ce.	92
,, ..	Chromium	Cr.	52.5
,, ..	Cobalt..	Co.	59
,, ..	Copper	Cu.	63.5
,, ..	*Didymium*	Di.	96

TABLE OF ELEMENTARY SUBSTANCES—*continued.*

EQUIVALENT POWER.	NAMES OF THE ELEMENTS.	SYMBOLS.	COMBINING NUMBERS.
Diatomic ..	*Glucinum*	G.	14
" ..	Iron	Fe.	56
" ..	*Lanthanum*	Ln.	92.8
" ..	Lead	Pb.	207
" ..	Magnesium	Mg.	24
" ..	Manganese	Mn.	55
" ..	Mercury (Hydrargyrum)	Hg.	200
" ..	Nickel	Ni.	59
" ..	*Palladium*	Pd.	106.5
" ..	Strontium	Sr.	87.5
" ..	*Tellurium*	Te.	129
" ..	*Thorium*	Th.	119
" ..	*Uranium*	U.	120
" ..	Zinc	Zn.	65.2
Triatomic ..	Aluminum	Al.	27.4
" ..	Antimony (Stibium) ..	Sb.	122
" ..	Arsenic	As.	75
" ..	Bismuth	Bi.	210
" ..	Gold (Aurum)	Au.	197
" ..	*Rhodium*	R.	104.2
Tetratomic ..	*Niobium*	Nb.	97.5
" ..	Platinum	Pt.	197.5
" ..	*Tantalum*	Ta.	137.6
" ..	Tin (Stannum)	Sn.	118
" ..	*Titanium*	Ti.	50
" ..	*Zirconium*	Zr.	89.6
Undetermined	*Erbium*	E.	?
"	*Indium*	In.	71.8
"	*Iridium*	Ir.	198
"	*Molybdenum*	Mo.	96
"	*Norium*	No.	?
"	*Osmium*	Os.	199
"	*Ruthenium*	Ru.	104.2

TABLE OF ELEMENTARY SUBSTANCES—*continued.*

EQUIVALENT POWER.	NAMES OF THE ELEMENTS.	SYMBOLS.	COMBINING NUMBERS.
Undetermined	*Terbium*	Tb.	?
„	*Tungsten* (Wolframium)	W.	184
„	*Vanadium*	V.	137.2
„	*Yttrium*	Y.	68

7. All the elemementary bodies may be regarded as made up of small particles or atoms, differing from each other in *weight.*

8. The atom is to be regarded as incapable of existing in a separate state, so that whenever a substance is spoken of as uncombined, or in a free state, a *double atom*, or *molecule*, is intended.

Thus, though one atom of Chlorine (Cl) may unite with one atom of Sodium (Na) to form Chloride of Sodium (NaCl), or common salt, Chlorine when spoken of as in a free state, is to be treated as a molecule or double atom (ClCl).

Similarly, free Hydrogen will be symbolized, as HH; free Oxygen as OO; and so on.

9. The numbers in the fourth column of the preceding table, represent the proportions by weight in which the different elements combine with each other, and are, therefore, termed *Combining proportions*, or *Combining weights.*

10. As these numbers are considered to represent the proportionate weights of an atom, or equal volume of each element, they have been also termed *Atomic weights.*

The term *equivalent* was, till lately, used as synonymous with atomic weight or combining proportion, as it was believed that bodies replaced each other in chemical combinations in the proportion of their atomic weights. (Chap. xvii. § 3.)

11. Hydrogen being the lightest of the elements, is taken as the standard of weight and volume, and all other elements are compared with it.

12. Since a given volume of Oxygen is 16 times as heavy as an equal volume of Hydrogen, the atomic weights, or combining proportions, of Oxygen and Hydrogen are taken respectively as 16 and 1.

Similarly, since Chlorine is $35\frac{1}{2}$ times as heavy as an equal volume of Hydrogen—Nitrogen, 14 times—Carbon vapour 12 times—Sulphur vapour (at 1900° F.), 32 times—the numbers 35.5, 14, 12, 32 are taken as the atomic weights or combining proportions of Chlorine, Nitrogen, Carbon and Sulphur, respectively.

13. The volume or space which the combining weight of a gas occupies, is called the *Atomic*, or *Combining Volume*, of that gas. (Chap. ix. § 17, &c.)

14. The symbol of each element is a short way of writing its name, and represents one atom or combining proportion of it.

15. The *symbol of a compound* is formed by writing together the symbols of the elements which compose it.

16. The combining weight of a compound is obtained by adding together the combining weights of its component parts; thus:—

Common Salt, or $NaCl = 58.5$; for $Na + Cl = 23 + 35.5 = 58.5$.

Caustic Potass, or $KHO = 56.1$; for $K + H + O = 39.1 + 1 + 16 = 56.1$.

Quick Lime, or $CaO = 56$; for $Ca + O = 40 + 16 = 56$.

Vermilion, or $HgS = 232$; for $Hg + S = 200 + 32 = 232$.

17. If, in writing the symbol of a compound body, more than one atom of an element has to be expressed, it is done by placing a *small* figure to the *right* of and below the symbol; thus:—

Water, or $H_2O = 18$; for $2H + O = 2 + 16 = 18$.

Chalk, or $CaCO_3 = 100$; for $Ca + C + 3O = 40 + 12 + 48 = 100$.

Saltpetre, or $KNO_3 = 101.1$; for $K + N + 3O = 39.1 + 14 + 48 = 101.1$.

Sulphuric Acid, or $H_2SO_4 = 98$; for $2H + S + 4O = 2 + 32 + 64 = 98$.

Lunar-caustic, or $AgNO^3 = 170$; for $Ag + N + 3O = 108 + 14 + 48 = 170$.

18. A large figure placed to the left of a group of two or more symbols, multiplies all the symbols which follow it, so far as the first comma, full stop, or sign of addition; thus :—

$3H_2O$ or $3(H_2O)$ means 3 atoms of water, or $H_2O + H_2O + H_2O$, $= 6H + 3O$.

$2KNO_3$ or $2(KNO_3)$ means 2 atoms of nitrate of potassium, or saltpetre, or $KNO_3 + KNO_3 = 2K + 2N + 6O$.

Sugar of Lead is composed of 1 atom of lead and 2 atoms of acetic acid, and may be expressed by

$$Pb\ 2(C_2H_3O_2),\ \textit{i.e.}\ Pb + 4C + 6H + 4O.$$

19. Combination between bodies themselves compound is often indicated by a comma or sign of addition placed between them; thus :—

Red Lead may be written $2\ PbO,\ PbO_2$, or more strictly, Pb_3O_4.

Common Alum, $Al_2 3SO_4,\ K_2SO_4 + 24H_2O$.

Malachite, $CuO,\ H_2O + CuCO_3$.

20. A collection of chemical symbols like the foregoing examples, is called a *formula;* and such a formula may be either *empirical* or *rational*.

21. An empirical formula is a simple enumeration of the elements composing a compound body, and the proportions in which they occur.

22. A rational formula attempts to show by the particular arrangement of the symbols, the mode in which the body is believed to be composed.

Thus, the empiric formula for slaked lime is CaH_2O_2, but the rational formula $CaO,\ H_2O$ shows that it consists of Lime (CaO), and Water (H_2O), and that it is a Hydrate of Lime.

Again, the empiric formula for Red Lead is Pb_3O_4, for it consists of 3 atoms of Lead and 4 of Oxygen; but a rational formula for it would be ($2PbO,\ PbO_2$), which expresses that it consists of 2 atoms of the protoxide of lead combined with 1 atom of the peroxide.

N.B.—A body can have but one empirical formula, but

it may be represented by various rational formulæ according to the different ideas entertained as to the arrangement of its component parts.

EXERCISES ON CHAPTER II.

1. How many elements are there, and how may they be arranged?
2. What elements form the group of metalloids?
3. What elements are included in the second group?
4. What is meant by the letters in the second column of the list?
5. How are the symbols formed, and from what are they sometimes derived?
6. How do you account for the names of some of the elements?
7. Of what do the elementary bodies consist?
8. What is meant by a *molecule* of a body?
9. What is meant by *combining proportions* or *weights?*
10. Why are these numbers called *atomic weights?*
11. What element is taken as the standard of weight and volume?
12. How are the atomic weights, or combining proportions ascertained?
13. What is meant by the atomic volume of a gaseous body?
14. What does the symbol of each element mean?
15. How is the symbol of a compound body written?
16. How is the combining weight of a compound found?
17. How is more than one atom to be expressed?
18. What effect has a figure to the left of a group of symbols?
19. How is combination between compounds sometimes indicated?
20. What is a chemical formula?
21. What is meant by an "empiric formula"?
22. When is a formula said to be "rational"?

CHAPTER III.

CHEMICAL AFFINITY.

1. The sixty-five elements enumerated in the preceding chapter have been called the Alphabet of the Science of Chemistry, for as the letters of the alphabet may be joined together to form words, so these elements may be united to form compounds. (**Synthesis.**)

2. Each word in a language is made up of two or more letters, into which it can be separated; (**Analysis**); but the letters themselves cannot be divided. (**Elements.**)

3. Some letters occur more frequently than others; for instance, we find one or more of the vowels, a, e, i, o, u, and y, in every word; while the consonants, j, q, x, and z, are seldom met with. So it is with the chemical elements; for Oxygen, Hydrogen, Nitrogen, Carbon, and a few more, are found in a large number of compounds; whilst Selenium, Cerium, Titanium, &c., occur but rarely.

4. The power or force by which elementary bodies are united in order to form compounds, is termed *Chemical attraction*, or *affinity*.

5. Chemical affinity may be defined as, that power by which particles of *different* kinds of matter combine to form a compound *always possessing some properties very different from those of its component elements.*

6. Chemical affinity differs from the attraction of gravitation, in not acting on *masses* of matter at *sensible distances*, but only on the *particles* of matter in *close contact.*

7. Chemical affinity differs from the attraction of cohesion in uniting particles of *different* kinds of matter, while the attraction of cohesion keeps together only particles of the *same* kind.

8. Although none but dissimilar bodies unite by chemical

affinity, mere dissimilarity will not of itself ensure u thus, sand and sawdust, salt and sugar, oil and v though very different bodies, cannot be made to u hence, they are said to have no affinity for each other.

9. Bodies which have no affinity for each other m mixed together, but their mixture is merely mechanical no change of properties takes place.

Ex. a.—Put into a glass some oil and water; however much may be stirred or shaken together, the mixture will be simply *me ical*, as will be evident by their soon separating when the agi ceases, the oil floating on the top of the water as at first.

Ex. b.—Add to the oil and water a little pearlash or soda, sti mixture as before, and the three will unite *chemically*, forming lution of soap.

Ex. c.—To the solution thus formed add a little diluted sulp acid. The acid will decompose the soap, and unite with the pea or the soda, while the separated oil will once more float on the su of the liquid.

10. From the last experiment we learn that chemical traction acts among different bodies with different deg of force.

11. The simplest cases of chemical affinity are thos which two elements unite to form a *binary* compound: is, one which contains only two different kinds of matte

Ex. d.—Cut two or three small pieces of phosphorus (P), place t on a tile or plate, and when dry throw a little iodine (I) upon th The two elements will combine, heat will be given out, part of phosphorus will be inflamed, part of the iodine will go off in vap and the remainder will form iodide of phosphorus. (PI_3).

Ex. e.—Rub together in a mortar, for some time, two parts weight of sulphur (S) and one of mercury (Hg). The yellow co of the one, and the brilliant metallic lustre of the other, will di pear; the two elements having combined to form a black powder (phuret of mercury) formerly called Æthiops mineral (Hg S.)

If the sulphur be melted in a ladle and the mercury added, the s black substance will be formed, only in less time.

Ex. f.—Mix together some tinfoil and a globule or two of merc The two metallic elements will unite and form a pasty mass or amalg si[illegible] which is at the back of a looking-glass.

12. The power by which two elements combine has been termed Simple or Single Affinity; the same power may also be exerted between two compounds, and *in exceptional cases*, between an elementary and a compound substance.

Thus sulphur (S) and iron (Fe), each an element, readily unite, when heated, forming sulphuret of iron, Fe S; sulphuric acid (H_2SO_4) and oxide of iron (Fe O), each a compound body, combine with ease to form sulphate of Iron (Fe SO_4) :—but the simple body, sulphur, does not combine with the compound, oxide of iron, neither does pure iron unite with sulphuric acid.

13. Hence the important rule which holds good, in inorganic chemistry, with but few exceptions, that *elements combine only with elements, and compounds only with compounds.*

14. It not unfrequently happens that when two compound bodies are mixed together, each is decomposed, and two new compounds are formed totally distinct from the former. This has been termed Double decomposition.

Ex. g.—Into a solution of acetate of lead, consisting of acetic acid and oxide of lead, pour some solution of chromate of potassium, which consists of chromic acid and potassium. A beautiful yellow powder will be formed, which is chromate of lead or chrome yellow, while acetate of potassium will remain in solution.

Here we have an interchange of the constituent parts of the two solutions, as shown in the following diagram

Acetate of Potassium.

Acetate of Lead.	Acetic Acid.	Potassium.	Chromate of Potassium
	Lead.	Chromic Acid.	

Chromate of Lead.

The same decomposition may be thus expressed by symbols :—

$$\text{Pb } \bar{\text{A}}_2 + \text{K}_2\text{Cr O}_4 = 2\text{K}\bar{\text{A}} + \text{Pb Cr O}_4.$$

Ex. h.—Into a solution of chloride of mercury pour some solution of iodide of potassium; a beautiful vermilion precipitate is obtained

affinity, mere dissimilarity will not of itself ensure union: thus, sand and sawdust, salt and sugar, oil and water, though very different bodies, cannot be made to unite; hence, they are said to have no affinity for each other.

9. Bodies which have no affinity for each other may be mixed together, but their mixture is merely mechanical, and no change of properties takes place.

Ex. a.—Put into a glass some oil and water; however much they may be stirred or shaken together, the mixture will be simply *mechanical*, as will be evident by their soon separating when the agitation ceases, the oil floating on the top of the water as at first.

Ex. b.—Add to the oil and water a little pearlash or soda, stir the mixture as before, and the three will unite *chemically*, forming a solution of soap.

Ex. c.—To the solution thus formed add a little diluted sulphuric acid. The acid will decompose the soap, and unite with the pearlash or the soda, while the separated oil will once more float on the surface of the liquid.

10. From the last experiment we learn that chemical attraction acts among different bodies with different degrees of force.

11. The simplest cases of chemical affinity are those in which two elements unite to form a *binary* compound: that is, one which contains only two different kinds of matter.

Ex. d.—Cut two or three small pieces of phosphorus (P), place them on a tile or plate, and when dry throw a little iodine (I) upon them. The two elements will combine, heat will be given out, part of the phosphorus will be inflamed, part of the iodine will go off in vapour, and the remainder will form iodide of phosphorus. (PI_3).

Ex. e.—Rub together in a mortar, for some time, two parts by weight of sulphur (S) and one of mercury (Hg). The yellow colour of the one, and the brilliant metallic lustre of the other, will disappear; the two elements having combined to form a black powder (sulphuret of mercury) formerly called Æthiops mineral (Hg S.)

If the sulphur be melted in a ladle and the mercury added, the same black substance will be formed, only in less time.

Ex. f.—Mix together some tinfoil and a globule or two of mercury. The two metallic elements will unite and form a pasty mass or amalgam, similar to that which is at the back of a looking-glass.

12. The power by which two elements combine has been termed Simple or Single Affinity; the same power may also be exerted between two compounds, and *in exceptional cases*, between an elementary and a compound substance.

Thus sulphur (S) and iron (Fe), each an element, readily unite, when heated, forming sulphuret of iron, Fe S; sulphuric acid (H_2SO_4) and oxide of iron (Fe O), each a compound body, combine with ease to form sulphate of Iron (Fe SO_4) :—but the simple body, sulphur, does not combine with the compound, oxide of iron, neither does pure iron unite with sulphuric acid.

13. Hence the important rule which holds good, in inorganic chemistry, with but few exceptions, that *elements combine only with elements, and compounds only with compounds.*

14. It not unfrequently happens that when two compound bodies are mixed together, each is decomposed, and two new compounds are formed totally distinct from the former. This has been termed Double decomposition.

Ex. g.—Into a solution of acetate of lead, consisting of acetic acid and oxide of lead, pour some solution of chromate of potassium, which consists of chromic acid and potassium. A beautiful yellow powder will be formed, which is chromate of lead or chrome yellow, while acetate of potassium will remain in solution.

Here we have an interchange of the constituent parts of the two solutions, as shown in the following diagram

Acetate of Potassium.

Acetate of Lead.	Acetic Acid.	Potassium.	**Chromate of Potassium.**
	Lead.	Chromic Acid.	

Chromate of Lead.

The same decomposition may be thus expressed by symbols :—

$$Pb\ \bar{A}_2 + K_2Cr\ O_4 = 2K\bar{A} + Pb\ Cr\ O_4.$$

Ex. h.—Into a solution of chloride of mercury pour some solution of iodide of potassium; a beautiful vermilion precipitate is obtained.

The chlorine of the chloride of mercury has united with the potassium of the iodide of potassium, while the iodine has combined with the mercury, forming the beautifully coloured iodide of mercury.

Chloride of Potassium.

Iodide of Potassium.	**Potassium.**	**Chlorine.**	**Chloride of Mercury.**
	Iodine.	**Mercury.**	

Iodide of Mercury.

Or symbolically, thus :—

$$2KI + HgCl_2 = 2KCl + HgI_2.$$

15. From the foregoing experiments, it will be seen that chemical affinity alters not merely the nature of bodies, but also their form and color. It likewise produces change of temperature, of volume, of taste, of smell.

Ex. i.—Put a drop or two of hydrochloric acid (HCl) into a small glass jar or wine-glass, and a drop or two of ammonia (NH_3) into another; invert one glass over the other, the two invisible gases emitted will unite, forming white fumes of chloride of ammonium, which will in a little time settle on the sides of the glass in a solid form. Thus two gases form a solid.

Ex. k.—Put a little hydrochloric acid (HCl) into a glass, and a few small pieces of chalk ($CaCO_3$). A brisk effervescence ensues, for a portion of the solid chalk (CO_2) passes off in the form of a gas.

Thus chemical affinity has power to change a solid into a gas.

Ex. l.—Pour into another glass some of the solution formed in the last experiment,—chloride of calcium ($CaCl_2$)—and add a few drops of sulphuric acid (H_2SO_4) ; a white solid is produced, which is sulphate of calcium,—plaster of Paris, or gypsum ($CaSO_4$).

Here we see two liquids have formed a solid.

Ex. m.—Put into a mortar equal weights of sulphate of sodium (Na_2SO_4) and nitrate of ammonium (NH_4NO_3). These two solid substances will, when rubbed together, gradually become fluid.

Ex. n.—Dissolve some loaf sugar in a *very small* quantity of hot water, so as to make a thick syrup, leaving some undissolved sugar at the bottom of the glass, which should be a somewhat tall one; add some *strong* sulphuric acid : the sugar will become black, being rapidly converted into charcoal. Sugar is composed of charcoal or carbon

and the elements of water; the strong acid unites with the water, and sets free a sufficient amount of heat to evaporate it, while the carbon is left behind.

In this experiment we have, first, a change of temperature; secondly, a liquid converted into vapor; and, thirdly, a liquid changed into a solid.

16. There are several causes which control chemical affinity, such as cohesion, heat, light, electricity, &c.

17. Since chemical affinity acts upon particles of different kinds of matter in close contact, the force of cohesion which unites particles of the same kind in masses, must be a hindrance to chemical action. Solid substances, even when finely powdered and mixed, very seldom combine chemically. Cohesion must, in most cases, be first overcome by dissolving one or both of the solids in a liquid, so that when the solutions are mixed, the particles may be brought into immediate contact.

Ex. o.—Mix together some tartaric acid and carbonate of soda (Na_2CO_3), *both quite dry*, no action will ensue; put some of the mixture into a glass of water, and effervescence arises immediately through the decomposition of the soda compound, and the escape of carbonic acid gas (CO_2).

18. Heat is generally the great opponent of affinity; since the latter draws the particles of matter together, while heat tends to drive them asunder, and by thus separating them from each other, often weakens and destroys their chemical union.

19. When the affinities which hold together the component elements of a compound body, have been destroyed by the application of a given degree of heat, other affinities come into action, and the elements arrange themselves into new combinations.

Thus all organic substances, as, wood, oil, tallow, &c., are decomposed by heat; but at the same time other compounds are formed, for no particle of matter can be destroyed.

20. Heat, by melting solid substances, and thus bring-

ing them into a liquid state, not unfrequently promotes chemical union.

Thus soda and sand may be mixed together without combining; but if the mixture be exposed to a red heat, it fuses, and then the sand and soda unite, forming a transparent glass.

21. Light has a powerful influence on the affinities between certain bodies. Hydrogen and chlorine gases will not unite if mixed together in the dark; but if brought into the sunshine they combine with explosion.

22. All photographic processes depend upon the chemical action exerted by certain of the rays of light upon various substances (generally compounds of silver) spread upon paper, glass, &c.

23. Electricity greatly influences affinity; indeed it has, by many persons, been supposed to be the foundation of all chemical action. Electrical currents from the galvanic battery are among the most powerful means of producing chemical decomposition.

Ex. p.—When two slips of platinum foil are dipped into a glass of water, to which a little sulphuric acid has been added, to make it a better conductor, and the slips connected by wires with the ends of a galvanic battery in action, the water is decomposed, bubbles of oxygen appearing at one slip (the positive), and hydrogen at the other (the negative), in the proportions of one, by measure, of the former gas to two of the latter.

24. Electro-plating, electro-gilding, and the electrotype process, are practical applications of the principle of chemical decomposition by the galvanic current.

25. Elements presented to each other in what is termed the *nascent** state, that is, at the moment they are set free from their previous combinations, often enter into new combinations which cannot be formed under other circumstances.

Thus hydrogen and nitrogen will not unite, though min

* *Nascent*, from *nasci*, to be born.

gled in the same vessel; but when these two gases are liberated together, by the decomposition of animal and vegetable matter, as in a manure heap, they readily combine to form ammonia (NH_3).

26. Sometimes a body by its mere *presence* or contact, *induces* changes in another, without itself undergoing any change. This *peculiar action* has been called *catalysis,** or *presence-action*. The reasons for it are not fully understood.

Thus in preparing oxygen from a mixture of chlorate of potass and manganese, the manganese suffers no change, and yet the gas is obtained much more easily than from the chlorate of potass alone.

Ex. q.—Mix a little starch with hot water, it will form a gelatinous or jelly-like mass, as used by laundresses. If some of this gelatinous starch be mixed in an evaporating basin with dilute sulphuric acid, and the mixture kept (by means of a lamp) *almost* at the boiling point for a few minutes, it will change into a semitransparent liquid. Now remove the lamp and neutralize the acid by stirring in some powdered chalk till all effervescence ceases; filter the liquid and evaporate. The starch has been changed into a kind of gum. This starch-gum is called *dextrine*.

If the starch and acid be boiled for some time, instead of only simmering, and the acid neutralized as before; on filtering and evaporating the liquid, a mass of sugar (grape sugar) will be formed. No portion of the sulphuric acid has been decomposed, neither has it suffered any gain or loss, for the sulphate of lime formed contains just as much acid as was originally employed.

27. Chemical affinity is governed by certain general laws, the first of which is, that *all chemical combination takes place in fixed, unchangeable proportions of weight.* (Definite Proportions.)

Thus, the water we drink is composed of two gases, Oxygen and Hydrogen: 16 grains, ounces, or pounds of Oxygen, unite with 2 grains, ounces, or pounds of Hydrogen, to form 18 grains, ounces, or pounds of water. If we attempt to form water by uniting these elements in different proportions, we shall fail, as a portion of one or other element will remain uncombined.

* *Catalysis*, from κατα-λυσις, an undoing.

28. The second law of Chemical affinity is, that *the same chemical compound always contains the same elements united in the same unvarying proportions.* **(Constancy of Composition.)**

Thus pure water at all places and at all seasons, contains the same proportions of its constituent gases.

29. The third law of Chemical affinity is, that *when elements unite to form more combinations than one, the larger quantities are multiples of the least by a whole number.* **(Multiple Proportions.)**

Thus Oxygen and Nitrogen form a series of five compounds (as below), which differ essentially in their properties; the amount of nitrogen in each is unchanged, but the quantity of Oxygen varies as the numbers 16, 32, 48, 64, 80, each number being an exact multiple of sixteen. The two gases will not combine in any other proportions.

Protoxide of Nitrogen	N_2O	$28+16 = 44.$
Deutoxide of Nitrogen	N_2O_2	$28+32 = 60.$
Nitrous Anhydride .	N_2O_3	$28+48 = 76.$
Hyponitric Anhydride	N_2O_4	$28+64 = 92.$
Nitric Anhydride . .	N_2O_5	$28+80 = 108.$

EXERCISES ON CHAPTER III.

1. Why may the elements be termed the A B C of chemistry?
2. Illustrate synthesis and analysis by the words of a language.
3. Name some of the more frequently occurring elements; also some of the rarer ones.
4. What is meant by chemical affinity?
5. What is the effect of chemical affinity?
6. How does chemical affinity differ from gravitation?
7. How does chemical affinity differ from attraction of cohesion?
8. When are bodies said to have no affinity for each other?
9. What is a mechanical mixture? Give illustrations of the difference between mechanical mixture and chemical combination.
10. Shew that there are different degrees of affinity.
11. What is meant by a binary compound?
12. What is simple or single affinity?
13. State, and illustrate, the fundamental rule of chemical union.

14. Explain and illustrate what is meant by double **decomposition.**
15. State and illustrate the changes produced by **chemical affinity.**
16. Name the causes which influence chemical action.
17. How does cohesion influence affinity ?
18. Why does heat often destroy chemical union ?
19. What becomes of the elements of a body that has been decomposed by heat ?
20. How does heat sometimes directly promote chemical union ?
21. How does light influence affinity ?
22. What is the principle on which photography depends ?
23. What influence has electricity upon chemical action ? Give an example.
24. On what principle does the art of electro-plating depend ?
25. Explain and illustrate what is meant by the nascent state of bodies.
26. What is meant by Presence-action, or Catalysis ?
27. State and explain the law of *definite proportions.*
28. What is meant by ot *constancy of composition ?*
29. Explain the law of *multiple proportions.*

CHAPTER IV.

METALLOIDS.

OXYGEN.

Symbol O. *Combining equivalent* 16.

1. **Oxygen** at the common temperature is a colourless, tasteless, inodorous gas. It was discovered by Dr. Priestly, in 1774, and by Scheele, a Swedish chemist, about the same time.

2. It was at first called dephlogisticated air. The term Oxygen was given to it by Lavoisier, a French chemist, who thought that the *acid* (Chap. vii.) properties of bodies were due to this element. The term Oxygen signifies *acid-producer.* (οξυς, sharp ; γενναω, I produce.)

3. Oxygen is the most widely distributed body in nature ; it forms one-fifth part by volume of all the air ; eight-

ninths, by weight, of all the water; one-half of all the solid crust of the globe; and more than one-half of all living plants and animals.

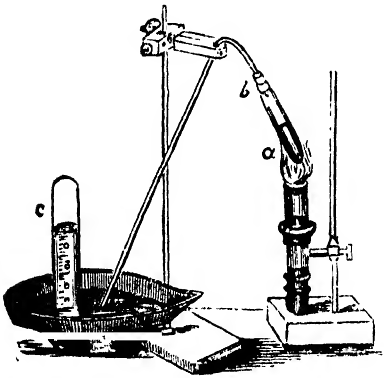


Fig. 1.

4. *Preparation.*—If a little red oxide of Mercury, HgO, be heated in a test tube *ab*, the affinity which holds the Oxygen to the Mercury is so feeble, that exposure to heat resolves the compound into oxygen gas and metallic mercury: thus, *Decomposition*:—$HgO = Hg + O$.

5. The gas may be collected in a jar *c*, over water at the pneumatic* trough,

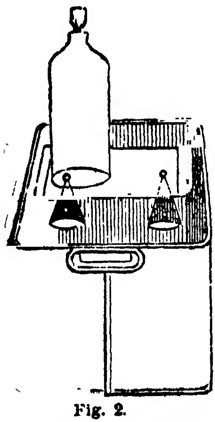


Fig. 2.

* "*The pneumatic trough*, fig. 2, is a tin box, 10 in. wide by 14 long. It is without a top, but has a shelf extending over a part of its surface, and rested on a small shallow ledge fastened to the sides of the box, so that the top of the shelf may be from 1 to 2 inches beneath the upper edge of the box, through the shelf are two or three holes, to which are attached on the under surface as many small funnels of tin as represented in the cut. Supposing gas is issuing out of the beak of a retort, and that beak were placed beneath the funnel, the shelf being covered with water; it is evident that the gas would pass through the funnelled hole and ascend into the air if the hole were uncovered, or else into a gas jar or other vessel placed over the hole. Also in passing through the water, the gas becomes purified and cooled. These are the objects of a pneumatic trough.

the metallic mercury condenses in globules on the cool part of the tube. A tolerably large tube should be used in this experiment. The arrangement is shown in the figure. This method of obtaining the gas is seldom adopted, owing to the expense of the oxide of mercury. It is interesting as being the mode in which it was first obtained by Priestly and Scheele.

6. A more economical method is by mixing about four parts by weight, of chlorate of potassium ($KClO_3$) with one part of the binoxide of manganese (MnO_2). These materials should be well mixed and dried, then introduced into a retort, as D, or into a Florence flask, furnished with a proper tube for delivering the gas, which must be collected over water at the pneumatic trough, A, as before. Heat should be applied *gradually*, since if this precaution be neglected the flask is likely to be broken. In a few minutes the gas comes off in great abundance. The jars intended to receive

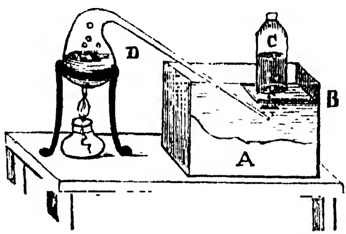


Fig. 4.

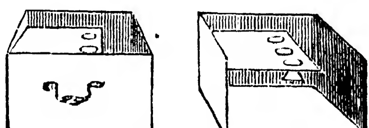


Fig. 3.

"Instead of tin, a common wooden box may be used. It is only necessary that it should be water-tight, and a common box may be easily made so, by running a little pitch or a mixture of resin and wax along the various joints—even a washing basin with a small three-legged stool standing in it will answer every purpose. A small foot tub makes an excellent trough, and a small garden pot, with a hole made in the side to receive the break of the retort, will make an equally excellent shelf to it, so easy is it to find chemical apparatus."—*Francis.*

the gas, as C, should stand filled with water, on the shelf of the trough B, *ready to be used* at once, as the gas generally comes over with great rapidity, and much would otherwise be wasted.

Decomposition :—$KClO_3 = KCl + 3O$.

7. The manganese in this case undergoes no change, it acts by mere contact (*catalysis*). The decomposition of the chlorate of potassium in the presence of the manganese takes place at a much lower temperature than if it were used alone, for then a heat would be required sufficient to soften, if not melt, the glass flask.

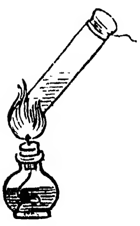


Fig. 5.

Ex. a.—Put a small quantity of chlorate of potassium into a test tube and hold it by a wire over a lamp. If when the chlorate is boiling and the gas is rising, a slip of burning wood be held above the chlorate and in the gas, the brilliancy of the flame will be increased; if the burning wood be allowed to fall into the melted chlorate, a sudden deflagration takes place.

8. Another method sometimes adopted for obtaining Oxygen, when absolute purity is not required, is to heat to redness in an iron bottle or gunbarrel some of the binoxide of manganese (MnO_2). The manganese oxide, at this high temperature, undergoes decomposition, but differs, however, from the red oxide of mercury in not yielding up the whole of its Oxygen.

9. In this decomposition, 3 atoms of the Binoxide of manganese (MnO_2) are resolved into one atom of the Protoxide of manganese (MnO); one atom of the Sesquioxide of manganese (Mn_2O_3); and a *molecule,* or double atom of Oxygen Gas, OO, thus :—

$$3MnO_2 = MnO + Mn_2O_3 + 2O.$$

10. The most striking property of Oxygen is its power of supporting combustion.

Ex. b.—Fasten a piece of candle or taper to a wire, light it and immerse it in a jar of oxygen gas; it will burn with much more than its usual brightness. If the taper be so extinguished, as to leave the wick red hot, and then immersed, it will be instantly relighted. The taper may be thus extinguished and lighted again several times with the same jar.

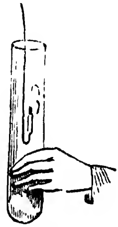


Fig. 6.

Ex. c.—Remove a jar of oxygen gas from the shelf of the pneumatic trough, by sliding it, with its mouth still under water, into a plate or saucer, and then lifting the whole together. Place in a deflagrating spoon, fig. 7, a piece of sulphur about the size of a pea, light it by the lamp, and immerse it in the jar of gas, taking out the stopper, which should have been well greased before-hand, to prevent its sticking. The sulphur will burn with a beautiful blue flame, and the jar will be filled with a vapour which is sulphurous acid (SO_2)* formed by the combination of the sulphur with the oxygen.

Ex. d.—Instead of sulphur introduce a piece of charcoal previously ignited at the spirit lamp. It will burn rapidly, forming carbonic acid (CO_2).* If the charcoal have the bark on, it will burn with very brilliant scintillations; fig. 8.


Fig. 7. Fig. 8.

Ex. e.—Put into a deflagrating spoon a piece of phosphorus about the size of a split pea, previously dried between the folds of blotting paper with as little handling as possible; ignite the phosphorus by touching it with a red hot wire, and immerse it quickly, but steadily, into a jar of Oxygen, to about two-thirds of the depth of the jar; fig. 9. The Phosphorus will burn with a light so intense that the eye can scarcely bear it; the jar will be filled

Fig. 9.

* Or, more correctly, Sulphurous and Carbonic anhydride.

with white clouds of Phosphoric anhydride (P_2O_5), which after some time will dissolve in the water.

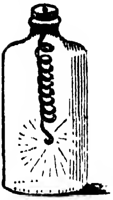


Fig. 10.

Ex. f.—A piece of iron wire (called binding wire) may be twisted into a spiral form like a cork-screw, and about the diameter of a lead pencil. The end of the spiral wire should be dipped in sulphur, or attached to a dry splinter of wood and ignited. When introduced into a jar of oxygen, the sulphur or wood kindles the iron, which burns with great brilliancy, while large melted drops of the black oxide of iron (Fe_3O_4) fall to the bottom of the jar.*

11. All the foregoing experiments are cases of combustion. Combustion is the combination of two or more bodies attended by light and heat; it is the result of intense chemical action.

12. In each of the above instances the light and heat are produced by the chemical union of oxygen with the burning body, the weight of which is increased exactly in proportion to the quantity of oxygen consumed.

13. Combustion in oxygen is the same in character as combustion in air; and the same products are formed. It is, however, much slower in the latter case because the oxygen of the atmosphere is diluted with four times its bulk of nitrogen.

14. Oxygen is the universal supporter of respiration, and is hence the great supporter of life. The lungs of land animals and the gills of fish are both adapted to the same

* When these experiments are given at the lecture table, the gas jar should not hold less than three pints, the phosphorus and sulphur should be introduced in a deflagrating spoon; the charcoal should be secured to a piece of wire, and the other end attached to a cork. The end of the iron spiral should also be attached to a cork, but there should be a piece of tin, or other metal, under the cork, to keep the latter from catching fire.

These experiments are best seen in a dark room. If smaller jars be used, the pieces of combustible material must be reduced in size.

purpose—to absorb oxygen; the former from the atmosphere, the latter from the water.

15. Pure oxygen, however, is not suited to sustain animal life; its effect is too powerful; over-excitement, fever, and in a short time death, would be the result of living in an atmosphere of undiluted oxygen.

16. If a living animal were introduced into a jar of oxygen, the circulation of the blood would be quickened, the respiration would be more frequent, the eyes would dilate, the blood vessels burst, and the poor creature die: it would, indeed, have been living too fast.

17. Oxygen gas is slightly soluble in water. This is of great importance to those plants and animals which live in water and are dependent on the dissolved oxygen of the water for the air they breathe.

18. Fishes die if placed in water deprived of its oxygen by boiling or otherwise. Hence, in an aquarium, it is necessary either to change the water frequently, or to provide a supply of living plants, which in their process of growth give out oxygen; this is absorbed by the water, and respired by the fish through their gills or breathing apparatus.

EXERCISES ON CHAPTER IV.

1. Give the symbol and combining weight of oxygen. By whom was oxygen discovered, and when?
2. Whence does oxygen derive its name?
3. Shew that oxygen is the most widely-diffused body in nature.
4. What is meant by HgO; and how may oxygen be obtained from it?
5. Why is the HgO process interesting? Why is it not usually adopted?
6. Explain the process of obtaining oxygen from a mixture of chlorate of potassium and binoxide of manganese.
7. What is the advantage of using binoxide of manganese with the chlorate of potassium?
8. How may oxygen be obtained from binoxide of manganese alone? In what respect does this process differ from that with the red oxide of mercury?

9. Give the formula for the decomposition of binoxide of manganese, and explain its meaning.

10. State and illustrate the most striking property of oxygen.

11. Define combustion.

12. How are the light and heat produced in combustion ?

13. In what respects does combustion in oxygen gas resemble, or differ from, combustion in common air ?

14. What relation has oxygen to animal life ?

15. Why would an atmosphere of pure oxygen be injurious ?

16. What would be the effect on an animal if it breathed only pure oxygen ?

17. Why is the solubility of oxygen in water important ?

18. Of what use are living plants in an aquarium ?

CHAPTER V.

HYDROGEN.

Symbol H. *Combining equivalent* 1.

1. **Hydrogen** is a gas at all known temperatures. It is colourless, tasteless, and without odour when pure, and the lightest body known, being 14 times lighter than common air. It cannot support combustion or respiration, but is itself highly inflammable, burning, however, only in the presence of oxygen.

2. Hydrogen was first obtained pure by Mr. Cavendish in 1766. Its name signifies *water-producer* (ὕδωρ, water ; γεννάω, I produce). It is, as its name implies, a constituent of water, as well as of all plants and animals.

3. Tallow, wax, gas, oil, coal, and all substances used either for heating or illuminating purposes, are **Hydrocarbons**, i. e. compounds of Hydrogen with another element **Carbon**.

4. Hydrogen is usually obtained by the decomposition

of water, H_2O. The simplest method of doing this, perhaps, is by passing a galvanic current through the water, whereby it is resolved into its constituent elements, the hydrogen appearing at the negative wire or electrode (—), and the oxygen at the positive (+); the quantity of the hydrogen being double that of the oxygen.

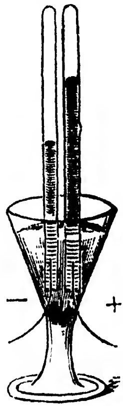


Fig. 11.

5. Another method of obtaining hydrogen, interesting from its simplicity, but too expensive to be used on a large scale, is by folding up a piece of potassium or sodium in blotting paper, and passing it quickly under a gas jar filled with water and standing mouth downwards on the shelf of the pneumatic trough, or into a test tube. (Fig. 11.) The metal decomposes part of the water, uniting with the oxygen to form potash, K_2O, or soda, Na_2O, which is dissolved in the remaining water; while the hydrogen rises to the top of the jar as a gas.

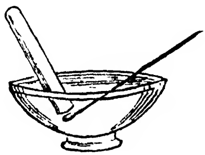


Fig. 11.

6. Decomposition : $K+H_2O = KHO+H$;

$$Na+H_2O = NaHO+H.$$

7. The presence of the dissolved potash or soda may be shewn by the restoration of the blue color to a piece of litmus paper previously reddened by an acid.

8. The presence of the gas may be shewn by removing the tube or jar from the water, having previously closed the mouth with the finger or a plate, inverting the jar, removing the covering, and applying a light, when the hydrogen will burn with a pale flame.

9. Iron at a red heat decomposes water, so if steam be made to pass from a flask B, through a red hot gun barrel *a b*, or porcelain tube filled with iron filings, it

will be decomposed into its elements, the oxygen uniting with the iron, while the hydrogen escapes through the tube *c d.*

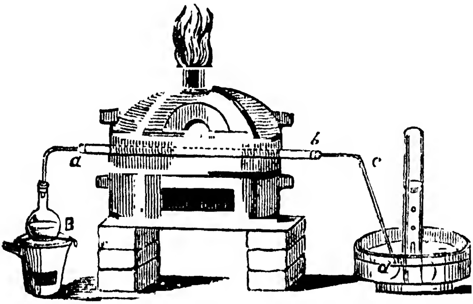


Fig. 12.

10. In this process four atoms of water are decomposed by three atoms of iron, forming one atom of black oxide of iron, and liberating 8 atoms of hydrogen gas.

11. The best method of preparing this gas is the following:—Take a wide-necked bottle fitted with a good sound cork, perforated with two holes, one to admit a tube for the escape of the gas, and the other for a funnel which should pass nearly to the bottom of the bottle. Put a few pieces of granulated zinc or some iron nails into the bottle, and pour in some dilute sulphuric acid through the funnel; gas will be rapidly given off, and may be collected at the pneumatic trough in jars filled with water.

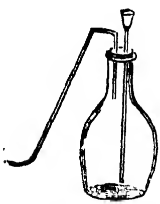


Fig. 13.

12. Care must be taken that the jars or other vessels in which the gas is collected contain no air, as hydrogen forms an explosive mixture either with oxygen or with common air.

For the same reason the first portions of the gas should always be allowed to escape, as they are invariably mixed with the air contained in the upper part of the bottle.

13. When zinc is placed in pure water, no chemical action is perceptible; but if a little sulphuric acid be added, decomposition of the water takes place, the oxygen of the water unites with the zinc, forming oxide of zinc (ZnO), which at once combines with the sulphuric acid, while the liberated hydrogen escapes as a gas.

14. *Decomposition:—*

$$H_2SO_4 + Zn = ZnSO_4 + 2H.$$
$$H_2SO_4 + Fe = FeSO_4 + 2H.$$

Ex. a.—Fill a small balloon made of gold-beater's skin with hydrogen; tie the mouth and let it free, it will immediately rise to the ceiling of the room; if out of doors, it will soon fly out of sight and be lost.

Ex. b.—Hydrogen is combustible, but not a supporter of combustion. This fact may be illustrated by introducing a lighted taper into a jar of hydrogen held with its mouth downwards: the hydrogen burns at the mouth of the jar, where it is in contact with air, but if the taper be moved farther into the jar its flame will be extinguished.

Ex. c.—Hydrogen gas may be ignited at a jet as it escapes from the bottle in which it is prepared, care being taken that *all the air is expelled* before the light is applied; it is as well also that the bottle should be loosely wrapped round with a cloth in case of any accident.

Ex. d.—If a dry gas jar be held over a jet of burning hydrogen, or over any ordinary flame, its interior becomes covered with drops of moisture. This moisture is due to the union of hydrogen and oxygen, forming water.

Ex. e.—If a bladder be filled with a mixture of two parts by measure of hydrogen and five of air, and a hole made in it with a pin, and a lighted taper applied, the mixed gases will explode with a loud report. If oxygen be used, two volumes of hydrogen should be taken and one of oxygen, being the proportions in which these gases unite to form water.

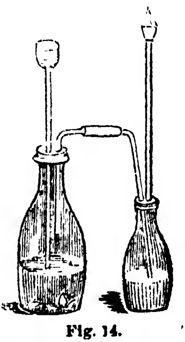


Fig. 14.

Ex. f.—Fill a gas bag or bladder provided with a stop-cock with the mixture of hydrogen and oxygen already mentioned; attach to it a tube or a tobacco-pipe bowl, and blow up some soap bubbles in a basin; *after* removing the tube, apply a lighted taper to the frothy surface: a violent explosion will ensue as before.

15. The flame produced by burning hydrogen is bluish or yellowish, and so faint that in daylight it is scarcely visible; it is, however, intensely hot, as may be seen by holding a piece of wire or a rod of glass to the end of the jet of burning gas.

16. The illuminating power of flame is almost entirely dependent upon the presence of solid matter in a state of combustion or intense ignition. In the hydrogen flame, we have two simple *gases* uniting, oxygen and hydrogen, giving rise to a *vaporous* compound, water, in the form of steam; hence we have but little light; in a dusty atmosphere, however, the flame becomes much more luminous, by igniting to intense whiteness the floating particles which come into contact with it.

Ex. g.—Shake some iron filings through the flame as it issues from the jet in *Ex. c.* They will be heated to whiteness, and form a shower of brilliant sparks. A piece of chalk sharpened to a point and held in the flame will emit a bright light. These experiments prove also the intense heat of the flame.

17. All ordinary flames burning in the air, or in pure oxygen, are of necessity hollow, combustion being only the energetic union of the substance burned with the surrounding oxygen, and this can evidently only take place in all common cases, at the surface of the burning body.

18. The flame of the mixed gases is, however, *solid*, for the combustible gas and the oxygen are already intermingled, and only require an increased temperature to cause them to combine in every part.

19. The flame thus produced is termed the oxy-hydrogen flame, and is still less luminous than that of hydrogen alone; its heat, however, is greatly increased.

20. **Various means have been employed for burning the explosive mixture of the two gases with safety.** Sometimes two gas holders or gas bags are used, one for each gas; the tubes from these reservoirs meet in another tube or jet, where the gases mix in very small quantity; fig. 15.

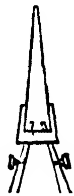


Fig. 15.

21. Sometimes a jet of peculiar construction called Hemming's jet, is employed, in which the gases, previously mixed in a bag, are forced through a tube filled with fine wires, which prevent the passage of the flame to the gas bag.

22. Another and more economical mode of using the two gases, and perfectly safe, is shewn in the diagram.

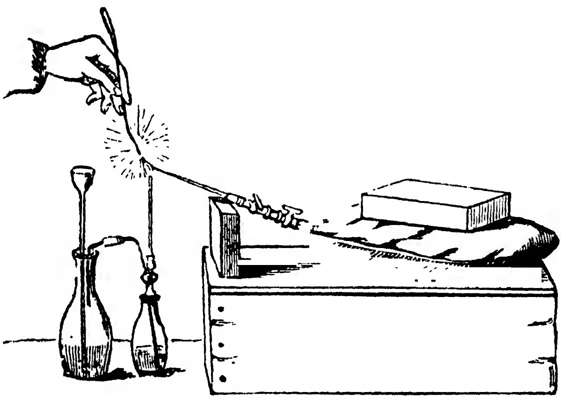


Fig 16.

The bag is filled with oxygen, which can be forced from it by the pressure of the weight; hydrogen is generated from zinc in the bottle, allowed as usual to pass through sulphuric acid in the second or drying bottle, and kindled at the jet. The oxygen is turned on, and brought within one eighth of an inch of the burning hydrogen, when the two gases burn together, and we obtain the oxy-hydrogen flame.

Ex. h.—Hold in the flame, by means of a pair of pliers, a fine piece of copper wire about an inch long: it will burn with a beautiful green flame.

Ex. i.—Instead of copper wire use a piece of platinum wire: it will burn with a white light. This metal cannot be burned, or even melted, by the heat of the most powerful furnace, yet it burns readily in the intense heat of the oxy-hydrogen flame.

Ex. k.—Treat a piece of watch-spring, or a small steel file, in the same way, it will burn with the greatest brilliancy, throwing off showers of sparks.

Ex. l.—Cut a piece of lime to a sharp point, and hold it in the flame; a light will be emitted nearly as dazzling as the sun. This is called the oxy-hydrogen or Drummond light, and has been used for a variety of purposes where intense light is required.

EXERCISES ON CHAPTER V.

1. Give the symbol and combining equivalent of hydrogen. State its principal properties.
2. Whence does hydrogen derive its name? By whom was it discovered?
3. Name its most useful compounds.
4. Whence is hydrogen obtained? Explain the mode of obtaining it by a galvanic current.
5. How can hydrogen be obtained by means of Potassium or Sodium? State the changes which take place.
6. Give the formulæ for the preceding decomposition.
7. How can the presence of the potash or soda be proved?
8. How may the presence of the hydrogen be shown?
9. What is the effect of passing steam over red-hot iron?
10. Write and explain the formula for the preceding decomposition.
11. What is the usual method of preparing hydrogen?
12. What special care is needed in collecting hydrogen?
13. Explain the usual process of obtaining hydrogen.
14. Give the formulæ and equivalents of the decomposition of water by zinc and iron.
15. What is peculiar with regard to the light and heat of inflamed hydrogen?
16 On what does the illuminating power of flame depend? How may this be proved? Why does burning hydrogen give but little light?
17. Why are ordinary flames hollow?
18. Why is the oxy-hydrogen flame solid?
19. What are the properties of the oxy-hydrogen flame?

20. How may the mixed gases be burned with safety?
21. Describe Hemming's jet.
22. Describe a more simple mode of safely burning the mixed gases

CHAPTER VI.

WATER.

Symbol H_2O; *combining weight* 18.

1. **Water**, or Oxide of Hydrogen, is a compound of two atoms of hydrogen and one atom of Oxygen; its atomic weight is consequently 18.

2. Pure water is a transparent, tasteless, inodorous liquid; colorless in small quantities, but in large quantities of a beautiful blue, as seen in the deep sea, and in the crevices of Alpine glaciers.

3. Water is 815 times heavier than air; a cubic foot of water at the temperature of 60°, weighs 1000 ounces (avoirdupois); or, since a gallon of water weighs 10lbs., we may say

A pint of pure water
Weighs a pound and a quarter.

4. Water readily assumes either the solid, liquid, or vaporous state, as ice, water, steam.

5. Under ordinary circumstances, water freezes at 32°, and boils at 212°.

6. Water, like other bodies, expands by heat and contracts by cold; but, unlike other bodies, after it has been cooled to a certain point (39°.3), it again expands.

7. A given weight of water occupies the smallest space, or is most dense at 39°.3; if its temperature vary either above or below 39°.3, it expands. This point is termed the point of greatest density of water.

8. While freezing, water expands very much, bursting the strongest vessels in which it may be contained; it is in this way that the hardest rocks are gradually crumbled down

into soil fit for vegetable life; for the moisture, which has during summer penetrated into the smallest cracks and hollows, freezes during winter, and by its expansive force breaks the solid stone.

9. Ice is lighter than water, and therefore floats on its surface; were it otherwise, so that the ice sank as fast as it was formed, rivers, lakes, and seas, would soon be converted into bodies of solid ice.

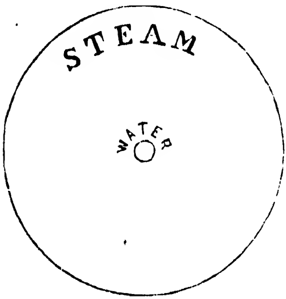


Fig. 17.

10. Steam occupies about 1700 times as much space as water; a cubic inch of water forms very nearly a cubic foot of steam.

11. Water has the power of dissolving many substances, and it is this property which renders it so important in nature and so serviceable to mankind.

12. Solution is a mere *mechanical* action, and does not produce any change in the *chemical* properties of a body; for on heating the solution, the water is driven off in vapor, and the particles of the dissolved body are recovered with their properties unaltered.

13. Transparency is essential to a solution. Where solid particles are distributed throughout a liquid, as when chalk is stirred with water, they are not dissolved, but only mechanically suspended in the water, and will separate by their own weight, or by the process of *filtering*.

14. Hot water has a greater solvent power for most substances than cold water. In the case of alum, for instance, hot water will dissolve twice as much as cold water. It follows that as the hot water becomes cold, part of the alum must again become solid. In so doing, the particles, in obedience to their mutual attraction arrange themselves

in particular forms or crystals; the process is termed *crystallization*. Sugar candy may be taken as another example.

15. In consequence of its great solvent power, water seldom, if ever, occurs in nature perfectly pure. Rain-water, as collected at a distance from large towns, and in the open fields, is the purest natural form of water, but even this contains a *trace* of ammonia, NH_3.

16. Absolutely pure water can only be obtained by the process of *distillation*,* whereby the water is converted into vapor (*evaporated*), which is afterwards condensed into the liquid state, while the substances that were previously dissolved in it are left behind in the retort or still.

In the chemist's laboratory and in many of the arts, large quantities of absolutely pure water are required. Its distillation is conducted in an apparatus represented in the annexed figure.

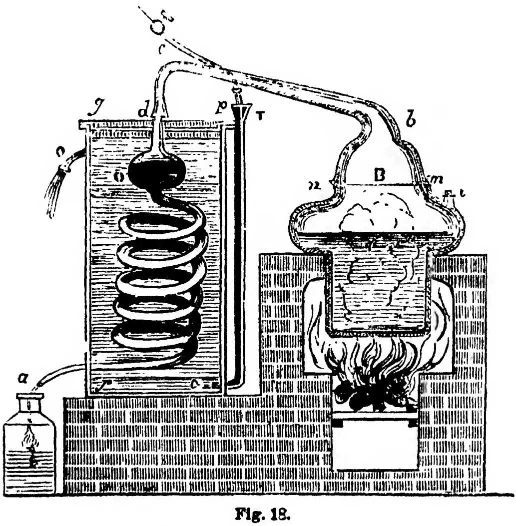


Fig. 18.

* *Distillatio* (Latin), falling in drops.

This distilling apparatus, or alembic, consists of a copper boiler, A, fixed in a brick furnace, having a dome-formed cover, B, adapted to it, from which a bent tube, *b c d*, proceeds, and is connected with a spiral tube called a *worm*. This worm is inclosed in a large cylindrical cistern, *p q j r*, constructed in metal, which is kept constantly filled with cold water. The lowest part of the worm passes out of this cistern near its bottom, and terminates at *a*, over the mouth of a jar, *c*, intended to receive the distilled water. An opening, *t*, having a steam-tight stopper, is provided in the boiler, through which the water to be distilled is introduced into it.

The vapor issuing from the boiler through the tube, *b c d*, passes into the worm, being first received by the vessel, o, where the condensation begins.

Passing next through the coils of the worm, it is exposed to the contact of its cold surface, and is entirely condensed and reduced to the liquid state before it arrives at the lower extremity, *a*, from which it trickles in drops into the jar, *c*.

17. Rain-water, as collected away from large towns, and water that springs from rocks, which, like granite or sandstone, are scarcely, if at all, soluble, is very pure, and will dissolve soap without curdling; it is called *soft* water.

18. Spring water, in general, however, contains lime and other substances, which it has dissolved out of the limestone and other strata through which it has flowed: such water is termed *hard* water; it will not dissolve soap, and is unfit for washing, &c.

19. Soap dissolved in spirits of wine is employed to test the hardness of water. It may be mixed with pure, soft water, without causing any change, but if added to a hard water, it produces a milkiness, the more considerable as the water is less pure.

20. Water unites *chemically* with other compound bodies forming a class of substances termed *hydrates* (ὕδωρ, water). Thus fresh-burned lime, CaO, unites rapidly with water in the process termed slaking, and is in reality converted into a hydrate of lime, CaO,H_2O.

21. Water is a product of all ordinary combustion. This has been shown in *Ex. d.* of the last chapter. It is also evident in cold weather on the windows of shops where

gas is burned, the vapor being condensed on the inner surface of the glass.

EXERCISES ON CHAPTER VI.

1. Give the symbol and equivalent of water.
2. What are the properties of pure water?
3. What is the comparative and what the absolute weight of water?
4. What different states does water assume?
5. Give the boiling and freezing points of water.
6. What is remarkable in the expansion and contraction of water by change of temperature?
7. What is meant by the *point of greatest density* of water?
8. What effects are produced by water when freezing?
9. What would happen were ice heavier than water?
10. State the comparative space occupied by steam and water.
11. What is the most important property of water?
12. What change does solution produce in the chemical properties of bodies?
13. When can water be purified by filtering?
14. Explain the process of crystallization.
15. What is the purest natural form of water?
16. How is absolutely pure water to be obtained?
17. Whence does soft water flow?
18. Why is some water termed hard water?
19. How may the hardness of water be detected?
20. Explain the term Hydrate.
21. Shew that water is a product of combustion.

CHAPTER VII.

COMPOUNDS OF THE METALLOIDS:—ACIDS, ETC.

1. Most of the non-metallic elements or metalloids combine with each other, the compounds so formed being either *acid* or *non-acid*.

2. The *non-acid* compounds of the metalloids with other elements have for the most part names ending in *-ide*; the termination *-uret* is more frequently used in some cases.

Thus the non-acid compounds of—

Oxygen are termed..	Oxides
Chlorine ,, .	Chlorides
Iodine are	Iodides
Bromine	Bromides
Fluorine	Fluorides
Carbon	Carbides or Carburets
Sulphur	Sulphides or Sulphurets
Phosphorus	Phosphides or Phosphurets.

3. An *acid* may be defined as a body which has a sour taste, reddens vegetable blues, and neutralizes, or destroys, the alkaline reaction of potass.

4. These are the *usual* characteristics of an acid, though they are not always found together; for some bodies, owing to their insolubility, have neither a sour taste, nor any action on vegetable colours, and yet must be regarded as acids, from their power of neutralizing potass or soda.

5. Thus common flint or sand, which is a compound of Oxygen and Silicon, if heated with potass (or soda) completely neutralizes its properties, and combines with it to form glass. Flint or sand has, therefore, been named

Silicic *acid* (SiO_2) though it neither tastes sour, nor acts on vegetable colours.

6. On the other hand, Carbonic acid (H_2CO_3), and Boracic acid (H_3BO_3), though they do not perfectly neutralize potass, have long been looked upon as acids, owing to their taste and action on vegetable blues.

7. It was formerly supposed that *all* acids contained Oxygen, and this supposition gave rise to the name *oxygen*, that is, *acid-producer*.

8. Many, however, of the compounds formed by Hydrogen with the other metalloids possess acid properties; they form the class of acids sometimes termed *Hydracids*, and their names begin with the prefix *Hydr-*, or *Hydro-*; as :—

Hydrochloric Acid	$HCl =$	36.5
Hydriodic „	$HI =$	128.
Hydrobromic „	$HBr =$	81.
Hydro-fluoric „	$HF =$	20.
Hydro-sulphuric	$H_2S =$	34

9. Oxygen unites with most of the other metalloids to form compounds which till recently were considered as acids, and formed the class of acids known as *Oxacids*.

10. Inasmuch, however, as these compounds have no acid properties *in themselves*, and manifest them *only when in combination with water*, they have no right to be called acids, and are now spoken of as **Anhydrides**.*

11. These Anhydrides, when in combination with water, form the class of acids formerly termed Oxacids, or *hydrated oxacids*.

12. The division into Hydracids and Oxacids may now be set aside, and all acids considered as compounds of

* When the water which was chemically combined with any substance has been removed, the substance is said to be anhydrous (*a*, *not*; and ὕδωρ, *water*); hence the term *Anhydride*.

Hydrogen, in fact, as Hydrogen-salts, in which the hydrogen acts as a base. (See Chap. xxvi.)

13. This will be seen by the following list of some of the most important of these Anhydrides, and their corresponding acids :—

	Anhydride.	Acid.
Carbonic	$CO_2 = 44$;	$H_2CO_3 = 62$.
Nitric	$N_2O_5 = 108$;	$HNO_3 = 63$.
Sulphurous	$SO_2 = 64$;	$H_2SO_3 = 82$.
Sulphuric	$SO_3 = 80$;	$H_2SO^4 = 98$.
Phosphoric	$P_2O_5 = 142$;	$H_3PO^4 = 98$.

14. When one element unites with another in several proportions, the different combinations are distinguished by suitable prefixes, marking the number of proportions of the *first-named* element which enter into union with one proportion of the latter.

15. A compound body, which contains only *one* proportion of each of its constituents, simply takes the name oxide, chloride, &c., as the case may be, according to § 2.

16. The Latin prefixes *bin-* or *bi-* (= twice), and *ter-* (= three times), are used when *two* or *three* proportions respectively of one body are combined with one proportion of another.

Thus we have binoxide of manganese $= MnO_2$; sesqui-oxide of gold $= Au_2O_3$; bichloride of mercury $= HgCl_2$, &c.

17. When two bodies combine in the proportion of *three* atoms of one body to *one* atom of the other, or as $1\frac{1}{2} : 1$, the compound takes the Latin prefix *sesqui*, which means one and a half.

Thus, common rust, Fe_2O_3 is termed sesqui-oxide of iron.

18. The Greek prefixes *proto-*, *deuto-*, *trito-*, &c., are properly used to denote respectively the first, second,

hird, &c., in a series of compounds formed by two elements, without reference to the proportions in which they occur.

19. The terms *proto*xide, *proto*-chloride, &c., are, however, frequently used simply in the sense of oxide, chloride, &c., as in § 15, instead of in their strictly grammatical sense of *first* oxide, &c. Any compound containing a *smaller* proportion of oxygen, chlorine, &c., would then be marked by the prefix *sub-* (under), as *Sub*oxide of copper, Cu_2O.

20. *Per-* or *hyper-* in a series of compounds, denotes the greatest though undefined quantity of the first element named, except when prefixed to the oxygen acids, when it implies the greatest amount of oxygen.

21. The termination *-ic*, in a series of oxygen acids, denotes the presence of more oxygen than the termination *-ous*; and *-ous* more than the prefix *hypo-* which means *under*; thus,—

$HClO$	= 52.5	form	*hypo*-chlor*ous*	acid.
$HClO_2$	= 68.5	,,	chlor*ous*	,,
$HClO_3$	= 84.5	,,	chlor*ic*	,,
$HClO_4$	= 100.5	,,	*per*-chlor*ic*	,,

EXERCISES ON CHAPTER VII.

1. How may the combinations formed by the metalloids with each other be arranged?

2. What compounds have names ending in *ide?* Give examples.

3. What are the usual characteristics of an acid?

4. Why is the usual definition of an acid too limited?

5. Why is flint, or silica, considered an acid?

6. Why should Carbonic and Boracic acids be regarded as acids?

7. What was formerly supposed to be an essential constituent of an acid?

8. What was intended by the term *hydr-acid?*

9. Explain the term *ox-acid.*

10. What are the compounds termed *anhydrides?*

11. How do the anhydrides become acids?

12. What are all acids at present considered to be?

13. Give the names and symbols of the principal acids.

14. How are different combinations of the same elements distinguished?

15. What is denoted by the terms oxide, chloride, &c.?

16. Explain and illustrate the use of the prefixes *bin-* or *bi-* and *ter-*.

17. Explain the prefix *sesqui-*, and give an example.

18. What do the Greek prefixes *proto-*, *deuto-*, *trito-* properly denote?

19. In what sense are the names having the prefix *proto-* often used? What is denoted by the prefix *sub-?*

20. Explain and illustrate the use of the prefixes *per-* and *hyper-*.

21. Explain the different prefixes and suffixes in a series of oxygen-acids.

CHAPTER VIII.

NITROGEN.

Symbol N. *Combining equivalent* 14.

1. **Nitrogen** is a permanently elastic fluid or gas, chiefly distinguished by its negative properties. It has neither colour, taste, nor smell; it neither burns nor supports combustion; it cannot maintain respiration, and it has no action upon vegetable colours. It is a little lighter than atmospheric air, and is slightly absorbable by water.

2. Nitrogen was discovered by Dr. Rutherford in 1772, and received its name (*nitre-producer*) from its connection with nitre or saltpetre, of which it forms a component part.

3. Lavoisier, the French chemist, who first investigated its principal properties, gave it the name Azote (*life-destroyer*), from the fact of its destroying animal life; this it does not by poisoning, but simply by excluding common air, and so causing suffocation.

4. Nitrogen is an important element, for it forms four-fifths of the air we breathe, and is a constant ingredient of all vegetables and animals. It occurs also in some minerals, as saltpetre (nitrate of potassium, KNO_3), cubical nitre (nitrate of sodium, $NaNO_3$), and a few varieties of coal.

5. Nitrogen may be prepared in several ways; but the easiest is by depriving atmospheric air of its oxygen, by burning some combustible, as sulphur, spirits of wine, or phosphorus, in a confined portion of air over water.

Ex. a.—Place a small piece of phosphorus, about the size of a pea, in a small porcelain capsule floating on the water in the pneumatic trough, or in a deep basin. Light the phosphorus, cover it with a jar of air, and let it burn out. The water will have risen in the jar and the space above will be filled with white fumes of phosphoric acid,

PO_5, these will, however, soon be absorbed by the water leaving the nitrogen pure.

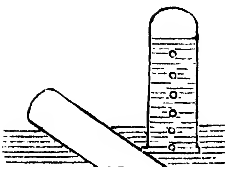


Fig. 19.

The gas so formed may be decanted into other jars, as in the diagram, and its properties examined.

Ex. b.—Attach a small piece of wax taper to a cork; light the taper, float the cork in a deep plate, full of water, and while the taper is burning cover it over with a glass receiver. It will be seen, that the light of the taper will become dimmed, and in a short time go out. This is because it has consumed all the oxygen of the air that was in the jar; what is left is nitrogen, mixed with some carbonic acid. It may be deprived of this latter by using lime water to stand the jar in, instead of common water. There must also be enough water that it may rise in the jar, according as the combustion goes on. The taper does not consume quite all the oxygen in the jar, but so nearly all that the nitrogen is sufficiently pure for ordinary experiments.

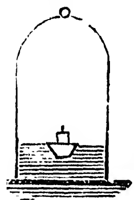
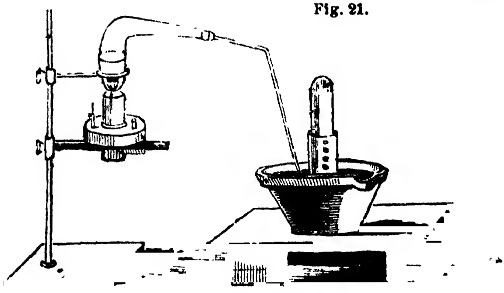


Fig. 20.

Ex. c.—If two jars be taken, one full of oxygen, and the other full of nitrogen gas, a lighted taper introduced into the nitrogen is immediately extinguished; but if it has previously been allowed to burn till the wick is red, and the least red spark remain on withdrawing it from the nitrogen,—it will be rekindled on transferring it quickly to the jar of oxygen gas, and again extinguished in the nitrogen; this may be repeated several times in the same portions of gas.

Ex. d.—Wash a piece of lean beef well, cut it into very small bits,

Fig. 21.



and put them into a glass flask or retort: pour in some diluted nitric acid, and then set the retort over a lamp, and receive the gas which comes over at the pneumatic trough in the usual way. It is nearly pure nitrogen.

6. Besides entering into the composition of atmospheric air, nitrogen combines with oxygen to form a series of five chemical compounds, remarkable as illustrating in a striking manner the law of multiple proportions (Chap. III. § 29) These compounds are as follow:—

Nitrous Oxide Gas	N_2O	=	44
Nitric Oxide Gas	NO	=	30
Nitrous Anhydride	N_2O_3	=	76
Hyponitric Anhydride	NO_2	=	46
Nitric Anhydride	N_2O_5	=	108

7. **Protoxide of Nitrogen, Nitrous Oxide. or Laughing Gas, N_2O.**—This is a colorless, transparent, odorous gas, having a faint, agreeable smell, and a sweetish taste. It is considerably heavier than common air. At 33° and under a pressure of 30 atmospheres* it becomes a liquid; and 150° below 0° it solidifies.

8. It supports the combustion of burning bodies with almost as much brilliancy as oxygen: from this, however, it may be distinguished by being readily absorbed by cold water, which dissolves ¾ths of its own volume; hence *warm water* must be used in collecting this gas at the pneumatic trough.

9. When taken into the lungs it produces in most persons a sort of intoxication, generally accompanied with pleasurable sensations, and not followed by unpleasant symptoms. *Sometimes* it acts injuriously, producing a sort of stupor, causing the experimenter to fall senseless for a few seconds, and then recover with confused ideas and headache: this, however, rarely happens, if the gas be *pure.*

* By an atmosphere is meant a pressure of fifteen pounds on each square inch of surface. Thirty atmospheres will therefore be equal to 450lbs. per square inch.

10. **Nitrous oxide** is prepared by placing solid nitrate of ammonium, (NH_4NO_3) in a retort, or in a flask with a bent tube, and gently heating it over a lamp The nitrate at first melts, and when the temperature rises to 340° F., it boils, and is entirely decomposed into water and gas; the latter may be received into jars, filled, as already mentioned, with warm water.

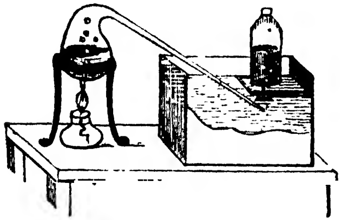


Fig. 22.

11. The following diagram, adapted from that given by Professor Graham, will illustrate the decomposition that takes places :—

	Before Decomposition.		After Decomposition.
Nitrate of Ammonium, $NH_4NO_3 = 80$.	Ammonia, $NH_3 = 17$	N = 14	Nitrous Oxide, $N_2O = 44$.
		H = 1	
		H = 1	
		H = 1	Water, $H_2O = 18$.
	Nitric Acid, $HNO_3 = 63$	H = 1	
		N = 14	
		O = 16	Water, $H_2O = 18$.
		O = 16	
		O = 16	
	80	80	80

Or in symbols :—

$$NH_4NO_3 = N_2O + 2H_2O.$$

Ex. 6.—Fill a wide test tube or small jar with the gas, and plunge

into it a taper with a glowing wick; the taper will be relighted and burn with brilliancy.

Ex. f. — Perform experiments with charcoal, phosphorus, and iron wire, similar to those described under oxygen (page 6), almost equally vivid combustion will ensue, and the products will be the same, for the burning bodies decompose the gas uniting with the oxygen and leaving the nitrogen in the jar.

Ex. g.—Perform a similar experiment with sulphur, and notice, that if the sulphur, when put into the gas, is only just burning, with a weak bluish flame, it will be extinguished; but if it is already burning vividly, its splendor will be greatly increased. Note also that the flame is of a beautiful rose color, and not blue as in pure oxygen.

Fig. 23.

Ex. h.—Fill a bag or large-sized bladder, fitted with a tube and stop-cock, with some of the gas, taking care that it is pure. Then closing the nostrils with one hand, *breathe out* as much air as possible from the lungs, apply the tube (with the stop-cock already turned on) to the mouth, and keeping it there, breathe the gas from and into the bag, as long as possible, and the intoxicating effect already described will be produced. This experiment should be performed in a large open space, lest the experimenter should injure himself or others, in his temporary madness.

It must never be tried by persons who have a determination of blood to the head.

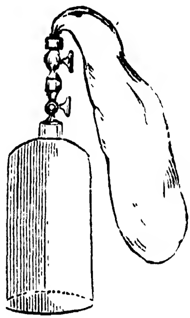


Fig. 24.

12. **Binoxide of Nitrogen or Nitric-oxide gas, N_2O_2 or NO.**—This gas may be prepared by pouring diluted nitric acid over some copper turnings or pieces of copper wire, in a retort or gas bottle, similar to that used in preparing hydrogen. No heat need be applied, and the gas may be received over cold water.

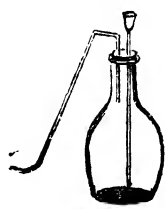


Fig. 25.

13. The gas is obtained by the deoxidation of Nitric acid (HNO_3), which is generally effected by means of copper (Cu).

For complete decomposition there must be eight atoms of the acid to three atoms of the copper, as may thus be shown:—

$$3Cu + 8(HNO_3) = 3(Cu2NO_3) + 4H_2O + 2NO.$$

The process may be divided into two stages, as follows:—

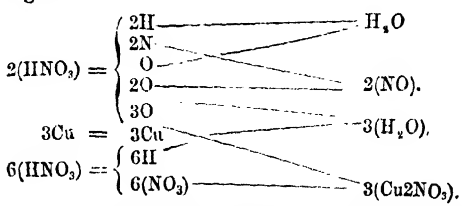
1st. Two atoms of the acid are acted on by the three atoms of the metal, each of which unites with one atom of oxygen, forming the protoxide of copper, CuO, while one atom of water (H_2O) and two atoms of Nitric oxide (NO) are liberated, thus;—

$$3Cu + 2(HNO_3) = 3CuO + H^2O + 2NO.$$

2nd. The three atoms of Oxide of Copper are dissolved in the remaining six atoms of acid; the three atoms of copper (a diatomic metal) replace the six atoms of hydrogen and form three atoms of nitrate of copper $3(Cu2NO_3)$; while three atoms of water $3(H_2O)$ are formed by the union of the hydrogen of the acid with the oxygen of the oxide; thus;—

$$3CuO + 6(HNO_3) = 3(Cu2NO_3) + 3H_2O.$$

14. The decomposition may also be shown by the following diagram:—



15. Nitric oxide is colourless and transparent, rather heavier than common air, very slightly soluble in water, incombustible, irrespirable, and incapable of maintaining ordinary combustion.

16. The most remarkable as well as most useful property of binoxide of nitrogen, is its power of instantly absorbing oxygen from the air, becoming then a dark reddish gas, hyponitric anhydride, $NO + O = NO_2$.

Ex. i.—Put some copper wire, or a leaf or two of Dutch gold (which is principally copper,) into a test tube, and add a few drops of nitric acid, the metal will be dissolved, and the upper part of the tube will be filled with orange red fumes of nitrous gas caused by the union of the binoxide of nitrogen with the oxygen of the air. If the tube be looked *into*, the solution appears green, but becomes blue on the addition of a little water: the green color is caused by the orange-colored gas mixing with the blue liquid; when this gas is dissolved by the water or otherwise removed, the blue color is made apparent

Ex. k.—Into a tall jar, half filled with nitrous oxide and standing on the shelf of the pneumatic trough, pass up a few bubbles of air, or, still better, of oxygen gas; notice the red fumes of nitrous acid (NO_2) produced, and the rise of the water, owing to the rapid absorption of the acid.

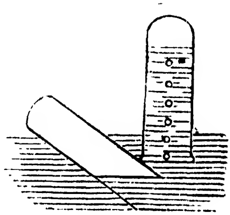


Fig. 26.

Ex. l. — Perform the same experiment, having first, either filled the jar with water colored blue by litmus, or put into it a piece of blue litmus paper; the reddening of the blue color proves the formation of an acid.

Ex. m.—Introduce a lighted taper or a piece of burning sulphur into a jar of nitrous oxide; the flame will be extinguished. The heat evolved being insufficient to decompose the gas.

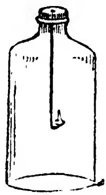


Fig. 27.

Ex. n. — Instead of sulphur use charcoal or phosphorus; these will burn with great brilliancy, the high temperature decomposing the gas, which yields its oxygen to the burning body at the same time that the nitrogen is disengaged.

17. **Nitrous Anhydride, N_2O_3.**—This substance may be prepared by mixing four measures of nitric oxide with one of oxygen, both perfectly dry, and exposing the mixture to great cold, by which it is condensed into a volatile green liquid.

It is of no practical application or importance.

18. **Hyponitric Anhydride, NO_2.**—This body is termed by some chemists **Hyponitric Acid.** It has already been noticed as forming the red vapours arising from the combination of nitric oxide with oxygen. By great cold it can be condensed into a very volatile and nearly colourless liquid.

Both this and the previously named body are decomposed by water, being resolved into binoxide of nitrogen and nitric acid.

19. Hyponitric Anhydride may be obtained by heating carefully dried nitrate of lead in a retort, and condensing the vapour in a receiver surrounded with ice.

$$4(Pb2NO_3) = 2PbO + 2O + 4NO_2.$$

20. **Nitric Acid,** HNO_3.—This is the most important of the chemical compounds of oxygen and nitrogen. It has been known from an early period, under the name *aqua fortis*, which it obtained from its power of rapidly dissolving most of the metals.

21. Nitric acid may be obtained by distilling in a retort equal weights of nitrate of potassium (KNO_3), and sulphuric acid (H_2SO_4), that is, one proportion of the former and one of the latter. The sulphuric acid displaces the nitric acid and unites with the potassium to form acid-sulphate of potassium, thus—

$$KNO_3 + H_2SO_4 = KHSO_4 + HNO_3.$$

22. Pure nitric acid is a colourless liquid (spec. grav 1.517). It is exceedingly corrosive, and has an intensely sour taste. It stains the skin and nails of a deep and permanent yellow, and rapidly decomposes animal and vegetable substances.

23. The acid thus described is in reality a *hydrated acid,* consisting, when strongest, of 1 molecule of absolute nitric acid N_2O_5, and 1 molecule of water H_2O, so that its symbol is $H_2O,N_2O_5 = H_2N_2O_6 = 2(HNO_3)$.

24. **Nitric Anhydride,** N_2O_5. — This substance has

been obtained by passing a current of dry chlorine gas, Cl, over dry and heated nitrate of silver, $AgNO_3$, thus—

$$2Cl + 2AgNO_3 = 2AgCl + O + N_2O_5.$$

It appears in the form of white prismatic crystals; though it has not yet been thoroughly examined, it seems to be destitute of acid properties so long as it continues in the anhydrous state.

25. Nitric acid is used for etching on copper, for dissolving and separating metals, also by surgeons as a caustic. In consequence of the large proportion of oxygen it contains and the facility with which it parts with it, nitric acid is frequently used by the chemist as a *means of oxidation.**

26. Nitric acid occurs in nature, chiefly in combination with potass, forming nitrate of potassium (KNO_3) or saltpetre; or with soda, forming nitrate of sodium ($NaNO_3$) or cubical nitre.

27. Nitric acid is found in small quantity in rain water, especially after thunderstorms. Every flash of lightning, and every electric spark which passes through the air, is believed to cause a minute portion of its oxygen and nitrogen to combine and form nitric acid, which is dissolved by the falling rain.

Ex. o.—Let nitric acid vapor pass slowly through a heated earthen-

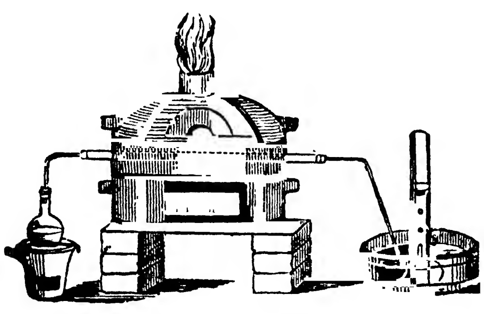


Fig. 28.

* When a substance combines with oxygen, it is said to be *oxidised* or to undergo *oxidation*.

ware tube: it will be decomposed into oxygen and nitrous acid if exposed to a dull red heat, and into oxygen and nitrogen gases if the tube be heated to whiteness.

Ex. p.—If a weak solution of indigo in sulphuric acid be poured into a glass and a little nitric acid be added, the blue color will be destroyed, especially if the liquid be heated.

Ex. q.—The acid properties of this substance will be proved by tasting it, *when greatly diluted,* and by its reddening litmus paper.

Ex. r.—Dip a quill, a little white worsted, flannel, or wool into some strong nitric acid, and it will become yellow, either directly or when exposed to the light for a short time. On removing the article from the acid it should be immediately well washed in water to prevent its destruction.

EXERCISES ON CHAPTER VIII.

1. State the properties of nitrogen.
2. By whom and when was nitrogen discovered? Why is it so named?
3. By whom was it termed azote, and why?
4. Why is nitrogen an important element?
5. State and explain some methods of obtaining nitrogen.
6. What remarkable series of compounds does nitrogen form with oxygen? Why is it remarkable?
7. Describe the properties of protoxide of nitrogen.
8. How may protoxide of nitrogen be distinguished from oxygen?
9. What is the effect of breathing nitrous oxide gas?
10. How may nitrous oxide gas be prepared?
11. Explain fully the decomposition of nitrate of ammonium by heat.
12. How may binoxide of nitrogen be procured?
13. Explain the decomposition in words.
14. State the same in symbols and by diagram.
15. Name some of the properties of nitric oxide gas.
16. What is its most useful property?
17. How may hyponitrous anhydride be prepared, and what are its properties?
18. What are the properties of nitrous anhydride?
19. How is nitrous anhydride prepared?
20. Give the symbol and equivalent of nitric acid.
21. How is nitric acid obtained? Give the formula of the process.
22. State the properties of nitric acid.
23. What is the real composition of nitric acid?
24. How has nitric anhydride been obtained, and what are its properties?
25. Name some of the uses of nitric acid.
26. How does nitric acid generally occur in nature?
27. How is the presence of nitric acid in rain water accounted for?

CHAPTER IX.

ATMOSPHERIC AIR.

1. **Air** is the transparent, colorless, invisible, and elastic fluid, which everywhere surrounds the globe, extending above its surface to the height of from 40 to 50 miles.

2. The air was formerly supposed to be a simple body, but is now known to be a mechanical mixture of several gases, the most abundant of which, are oxygen and nitrogen: besides these, it contains carbonic acid gas, and watery vapor, as well as traces of nitric acid and ammonia.

3. The quantity of these last named bodies present in the air, being dependent on local causes, varies considerably; not so with the oxygen and nitrogen. Air brought from great heights and great depths, as well as from regions of the globe far distant from each other, always presents these gases, combined in a uniform ratio.

4. Four-fifths of any volume of air are found to be nitrogen, and the remaining one-fifth, oxygen, or more exactly, 100 measures of air = 79·19 nitrogen + 20.81 oxygen.

5. Every 100 parts by *weight* of air contain 77 parts by weight of nitrogen and 23 of oxygen. The former of these numbers is not a multiple of 14, the atomic weight of nitrogen, neither is the latter a multiple of 16, the atomic weight of oxygen, as they would necessarily be, by the law of multiple proportions, if the air were a chemical compound, and not a mere mixture of the gases.

6. A better idea of the relative proportions of the elements of the air may be obtained by supposing the atmosphere to be of the same density throughout, and its elements arranged in strata according to their specific gravities. "In such a case, the atmosphere would be five miles in height. The greatest quantity of watery vapor would, if condensed, form a layer or stratum of water about five inches deep; the layer of carbonic acid, next above,

would be about thirteen feet deep; that of oxygen, one mile; and that of nitrogen, four miles."

7. The chemical properties of the air are chiefly those of the oxygen it contains. It is this gas, diluted and weakened in its energetic influence by four times its bulk of nitrogen, which renders atmospheric air the universal sustainer of animal and vegetable life.

8. The nitrogen in the air is not only useful, by diluting the oxygen as just stated, but also by increasing the bulk of the atmosphere without giving it any active chemical properties. "By this increase in the volume of the atmosphere, provision is made for the occurrence of winds, the diffusion of heat, the tempering of climate, and other useful ends essential to the welfare of man."

9. The atmosphere is the great conductor of sound; without it there would be perfect silence everywhere. It also refracts and reflects the rays of light; were there no atmosphere, there would be total darkness, except just at the spot on which the sun was shining.

10 The presence of oxygen in the air is shown by the products of combustible bodies when burned in it being exactly similar to their products when burned in pure oxygen. Thus, hydrogen burning in the air becomes water, the oxide of hydrogen; iron exposed to damp air becomes oxidised or rusted.

Ex. a.—Take a tube graduated or divided into a hundred equal parts, and place it mouth downwards on the shelf of the pneumatic trough. Fasten a piece of phosphorus to a stiff wire, and pass it up into the jar, and leave it so without disturbing it for 12 hours or more. The phosphorus will slowly combine with the oxygen of the air forming phosphoric acid, P_2O_5 while the water will rise in the tube, and on drawing out the phosphorus, will be found to occupy 21 divisions, leaving 79 divisions filled with nitrogen. This is the plan already described for obtaining nitrogen, only the decomposition in that case was more speedy in consequence of the ignition of the phosphorus.

11. If oxygen and nitrogen in their proper proportions of one to four be mingled in a receiver, we obtain an artificial air having the same properties as natural air. Air is thus shewn to be a mechanical mixture, and not a chemical

compound, for none of the properties of its constituent elements has been changed, and no diminution of volume has taken place, though one, if not both, of these results occur in every case of chemical combination.

12. The thorough intermingling of the gaseous elements, which causes the air to have everywhere the same uniform composition, is brought about by what is called the *law of gaseous diffusion*.

13. The operation of this law may be shewn thus :—

Ex. b.—Let two vessels be placed one above the other, and connected by a narrow tube of convenient length, fig. 29. Let the lower vessel be filled with carbonic acid gas, (see page 67, § 7), and the upper one with hydrogen. After a short time the carbonic acid, though twenty times heavier than the hydrogen, will be found to have ascended into the upper vessel; while hydrogen will have descended into the lower one,—a complete intermingling of the two gases having taken place, against the action of gravity.

Fig. 29.

14. All gases possess this property of diffusing into each other, although at different rates of velocity, depending on their density: the lighter the gas, the more rapid is the diffusion, and consequently, a larger volume of the lighter gas passes in one direction, than of the heavier gas in the other.*

15. The effect will be produced, even though a piece of bladder, sheet India rubber, or some similar barrier intervene.

Ex. c.—Take a piece of wide glass tube ten or twelve inches long, and close one end with a plug of plaster of Paris, about half an inch thick, and well dried. Fill the tube with hydrogen by displacement, *i. e.* by holding it over a jet of the gas from a bladder, or from the generating bottle, fig. 14; and then set it mouth downwards, upright in a glass of water. The gas will be so rarefied by its diffusive power, being greater than that of common air, that the water will rise in the tube

* "The replacing volumes of the gases, and consequently the velocities of their diffusion, are *inversely* as the square roots of their densities."—*Wilson.*

several inches above the level of the water in the glass; for the light hydrogen has escaped faster than the heavier air could descend through the plaster of Paris to supply its place.

16. The foul air produced by animal respiration as well as that arising from numerous other sources of contamination, is perpetually removed by diffusion, and the atmosphere is thus preserved pure and respirable.

17. As the space occupied by a given weight of a gas varies according to the temperature and pressure, it is necessary to have a fixed standard of comparison which is now usually expressed according to the metric system of weights and measures:—

Standard Temperature = 0° Centigrade, or 32° Fahr.
Standard Pressure = 760 millimètres of mercury.

18. At the above standard temperature and pressure
1 Litre of Hydrogen weighs nearly .0894 grammes.

Hence, under the same circumstances:—

1 Litre of	Oxygen, weighs	$16 \times .0894 = 1.430$	grammes.
,,	Nitrogen ,,	$14 \times .0894 = 1.251$	,,
,,	Chlorine ,,	$35.5 \times .0894 = 3.173$	,,

19. Sometimes the standard pressure and temperature are taken on the English scale:—

Standard Temperature = 60° F.
Standard Pressure = 30 inches of mercury.

20. At this standard temperature and pressure:—
1 grain of Hydrogen occupies $46\frac{2}{3}$ cubic inches;

while 16	grains of	Oxygen	severally occupy the same space.
14	,, ,,	Nitrogen	
$35\frac{1}{2}$	,, ,,	Chlorine	
12	,, ,,	Carbon (vapour) ..	
32	,, ,,	Sulphur (vapour at 1900°)	
		&c., &c.	

21. The atomic or combining volumes of the *elementary* bodies in a gaseous state (under the same circumstances of pressure and temperature) are thus seen to be equal to each other, and are often symbolically represented by a *single* square, thus: □.

22. The atomic or combining volume of a compound gas or vapour is found to be double that of the simple gases, hence it is symbolically represented by a double square, thus: $\square\square$.

23. From the foregoing we obtain the two following laws:—

1. The densities of the elements in the gaseous state are proportionate to their combining weights. (The density of Hydrogen = 1.000).
2. The densities of the compound gases are (for the most part) proportionate to half their combining weights.

Thus, the density of Steam,	$H_2O = 18 \div 2 = 9.$	
	Hydrochloric Acid,	$HCl = 36.5 \div 2 = 18.25.$
	Carbonic Acid,	$CO_2 = 44 \div 2 = 22.$
	Ammonia,	$NH_3 = 17 \div 2 = 8.5.$
	Nitric Oxide,	$NO = 30 \div 2 = 15.$
	Marsh Gas,	$CH_4 = 16 \div 2 = 8.$
	Olefiant Gas,	$C_2H_4 = 28 \div 2 = 14.$

24. The densities of the simple and compound gases thus obtained on the hydrogen scale (Hydrogen = 1.000), may be readily reduced to the usual or atmospheric scale (Air = 1.000) by multiplying them by .07, or more accurately by .0692, which is the density of hydrogen as compared with air.

25. All gases expand alike, or very nearly alike, when their temperature is raised.

This expansion is found to be $\frac{1}{273}$ of their volume at the freezing point (0° C.) for every increase in temperature of 1° Centigrade, or $\frac{1}{490}$ of their volume at the freezing point (32° F.) for every increase in temperature of 1° Fahrenheit.

26. Hence the altered volume of a gas at a different temperature and pressure may be readily calculated, thus:—

METRIC SCALE.

$$\text{As given pressure : Standard pressure} \quad \left.\begin{array}{r} \\ 273 : 273 + \text{no. of degrees} \\ \text{above } 0^\circ \text{ Cent.} \end{array}\right\} :: \left\{\begin{array}{c} \text{given vol.} \\ \text{at } 0^\circ\text{C.} \end{array}\right\} : \left\{\begin{array}{l} \text{altered} \\ \text{volume} \end{array}\right.$$

ENGLISH SCALE.

$$\text{As given pressure : Standard pressure} \quad \left.\begin{array}{r} \\ 490 : 490 + \text{no. of degrees} \\ \text{above } 32^\circ \text{ Fahr.} \end{array}\right\} :: \left\{\begin{array}{c} \text{given vol.} \\ \text{at } 32^\circ \\ \text{Fahr.} \end{array}\right\} : \left\{\begin{array}{l} \text{altered} \\ \text{volume} \end{array}\right.$$

EXERCISES ON CHAPTER IX

1. What is the air?
2. Of what is the air composed?
3. What is meant by the uniform composition of the air?
4. State the proportions *by volume* of oxygen and nitrogen in the air.
5. What are the proportions *by weight* of the constituent gases of the air? What may be inferred from these proportions?
6. Illustrate the relative proportions of watery vapor, &c. found in the air.
7. On what do the chemical properties of the air chiefly depend?
8. Of what use is the nitrogen in the air?
9. How does the atmosphere influence light and sound?
10. How may the presence of oxygen in the air be proved?
11. How may air be formed artificially? How may air be shewn to be a merely mechanical mixture?
12. How is the uniform composition of the air secured?
13. How may the diffusive power of different gases be illustrated?
14. State the law of gaseous diffusion.
15. Illustrate the different velocities with which gases diffuse into each other.
16. What benefits result from the diffusive power of gases?
17. Why is a standard temperature and pressure required?
18. State the weight of a litre of hydrogen, and calculate the weight of equal quantities of the other gases.
19. Give the standard temperature and pressure (i.) on the metric system; (ii.) on the English system.
20. What space is occupied by one grain of hydrogen?
21. How are the combining volumes of the gaseous elements symbolized?
22. What symbol expresses the combining volume of a compound gas?
23. What general laws may be stated relative to the densities of gaseous bodies?
24. How may the densities of gaseous bodies according to the hydrogen scale be reduced to the standard of atmospheric air?
25. In what ratio do gases expand by alteration of temperature?
26. How may the altered volume of a gas be calculated?

CHAPTER X.

CARBON.

Symbol C. *Combining equivalent* 12.

1. **Carbon** is an elementary non-metallic body, widely **diffused** throughout nature, both in a free state and in

combination with other elements. It is solid, tasteless, inodorous, and absolutely infusible by any degree of heat at present known.

2. The purest and at the same time rarest form in which carbon occurs, is that of the **Diamond**. The diamond is carbon in a crystallized state; it is highly transparent and brilliant, and the hardest substance known. Diamonds are found in Golconda, Brazil, Peru, the West Indies, and the Uralian mountains.

3. **Graphite**,* or, as it is sometimes called, plumbago, is another variety of nearly pure carbon, having a metallic appearance, and containing a very small proportion of oxide of iron in a state of mixture. It occurs in veins, mostly among primitive rocks, in Cumberland and various other parts of the world.

4. Graphite is also frequently termed black-lead, but it contains no lead whatever. On account of its power of sustaining great heat without change, crucibles are frequently made of it; but its best known use is for making drawing-pencils, for which only the finer and softer kinds are employed; the coarser varieties are used to give lustre to iron, as stoves, &c., or to lessen the friction of machinery, instead of oil or grease.

5. **Anthracite**† is a third form of nearly pure carbon. It is slowly combustible, but without flame, hence it has been called *blind* coal; another name for it is *glance* coal, from the German *glanz* (lustre), because of its shining surface.

6. There are several artificial varieties of charcoal employed in the arts, such as *wood-charcoal*, *animal-charcoal*, *coke*, and *lamp-black*.

7. Common **Wood charcoal** is made by piling logs of wood in a conical heap, covering them with turf and

* Graphite, from γραφω, *grapho*, I write.

† Anthracite, from ανθραξ, *anthrax* charcoal.

mould, and burning the mass slowly, small openings being left here and there to admit a little air.

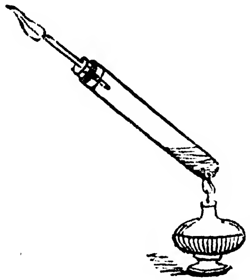


Fig. 30.

Ex. a.—Take a test tube fitted with a cork, and a piece of glass tube or tobacco-pipe, put into it a few chips of wood and heat them over a lamp, the gaseous constituents of the wood, oxygen and hydrogen, with a little carbon, will be driven off in the form of vapor of water and inflammable gas, which latter may be ignited, as in fig. 30, while the charcoal remains behind.

8. **Animal Charcoal**, called also bone-black, or ivory-black, is a form of carbon, combining much mineral matter, obtained by heating bones and other animal substances in a retort, so long as any volatile products are given off.

9. Animal substances consist of oxygen, hydrogen, nitrogen, and carbon, combined with various earthy materials; when exposed to a high temperature, out of contact with the air, the gases are given off in different states of combination, and the carbon remains mixed with the mineral matter, principally phosphate of lime. If, however, the animal substances be heated in contact with the air, the carbon also will be burned away, leaving only a white earthy mass behind, as may be seen by heating a bone in an open fire.

It is this earthy matter which gives to bones their hardness and rigidity.

10. **Coke** is the black, porous mass left after heating coal in retorts, with the air excluded, as is done in the manufacture of illuminating gas.

Coke ignites with difficulty, but gives out more heat when burning than an equal bulk of any other fuel.

11. **Lamp-black** is the soot or condensed smoke arising from the imperfect combustion of tar, resin, and other in-

flammable substances containing hydrogen and carbon, with a limited supply of air. It is well known that an oil or camphine lamp, turned up too high or having an insufficient draught of air, smokes; this smoke, unless carried away by currents of air, condenses, and falls all round in "blacks." Lamp-black is much used in the manufacture of black pigments, such as printers' ink, Indian ink, Frankfort black, &c.

12. Charcoal is a very indestructible substance, since at ordinary temperatures it is scarcely at all affected by air or moisture. Wooden stakes or piles are rendered more durable by slightly burning or charring their surface before driving them into the ground.

13. Charcoal possesses in a remarkable degree the power of absorbing gases, and retaining them in large quantities. This property of absorption varies with different kinds of charcoal.

Ex. b.—Fill a small jar or phial with the gas or fumes of ammonia, and place in it a piece of newly-burned charcoal, and close the jar. After a few hours the whole of the ammonia will have been absorbed—no odor remaining. Box-wood charcoal is found to absorb ninety times its bulk of ammoniacal gas in twenty-four hours.

14. Intimately connected with this property of absorbing gases is the power which charcoal possesses of removing offensive odors and checking putrefaction. It is a powerful *antiseptic.**

The decay of meat and vegetables may be retarded by packing them in charcoal.

By charring the insides of casks, water may be kept a long time in them without spoiling.

The disagreeable scent attaching to clothes and other articles, from damp sea voyages, diseases, peculiar medicines, occupations, &c., may be removed by wrapping them for a few hours with some pieces of animal charcoal. The same substance forms also a valuable tooth-powder.

* Antiseptic, from αντι, *anti*, against, and σηπειν, *se-pein*, to cor-

15. Animal charcoal, and especially that variety termed bone-black, possesses an extraordinary power of removing coloring matters of animal or vegetable origin from their solutions. It is largely used by sugar refiners, who deprive brown sugar of its color by boiling it with charcoal made from charred blood, or filtering it through a layer of that substance coarsely pounded.

Wood-charcoal has very little decolorizing power.

Ex. c.—Fold a piece of filtering paper or white blotting paper, so as to fit into a funnel. Fill it two-thirds full of bone-black, and let some water, colored with a few drops of black or red ink, filter through. The liquid will be rendered colorless.

Ex. d. — The charcoal acts more powerfully if heated with the liquid to be bleached. Boil common vinegar, some syrup of brown sugar, or some beer with a few spoonfuls of bone-black, all the color will be absorbed, and the liquid may be st ained off clear.

16. The affinity of carbon for oxygen at a high temperature is very great. It deprives most metallic oxides of their oxygen, and of course brings them into the metallic state. An agent which thus brings out metals from the ore, is called a *reducing* agent, and the process is termed *reduction.*

Ex. e.—Mix 4 parts of red lead with 1 part of charcoal powder. Put the mixture into a crucible, or tobacco-pipe bowl, and stir them well together, so that the color may be a dirty brown. Put the crucible into a clear fire, and give it a red heat for a quarter of an hour; when sufficiently heated, pour out the contents of the crucible, and metallic lead will run from under the powder at the top.

Ex. f.—If 2 grains of powdered charcoal and 4 grains of chlorate of potassium be carefully mixed in a piece of paper, then folded up, and placed upon an anvil—upon being struck by a hammer a violent detonation will take place.

N.B. It is by no means easy to pulverize charcoal when cold; therefore, when pounded charcoal is wanted, use a piece out of a charcoal fire, or burn a stick in an ordinary fire, cutting off the ignited portions.

17. Charcoal, on the other hand, appears to exert an oxidizing effect upon the gases and vapors it absorbs.

This effect is not *pro*duced but only *in*duced by the charcoal which condenses the gases within its pores, and brings them into closer contact with oxygen, causing them to lose their characteristic properties, and form new combinations.

Ammonia, for example, changes into nitric acid, and sulphuretted hydrogen into sulphuric acid. Thus :—

$$NH_3+4O = HNO_3+H_2O.$$
$$H_2S+4O = H_2SO_4.$$

18. Charcoal, saturated with ill-smelling and offensive matters, as already mentioned, may be advantageously used for manure, fertilizing the soil with which it is mixed, and gradually yielding them as valuable nourishment to growing plants.

19. Carbon affords a good example of one substance appearing under different forms and possessing different properties. It is sometimes *amorphous* or shapeless, as in coke; or crystallizing under two incompatible forms (*dimorphous*), as in the diamond and in graphite. The property, which several elements have, of thus assuming different forms is termed *allotropic*, from αλλοτροπος, changeable.

EXERCISES ON CHAPTER X.

1. State the symbol, combining equivalent, and general properties of carbon.
2. Describe the purest form of carbon.
3. What is graphite? and where is it found?
4. For what is graphite used?
5. What is anthracite?
6. Name some artificial varieties of carbon.
7. How is common wood-charcoal prepared?
8. What is animal charcoal?
9. Explain the process for obtaining animal charcoal.
10. What are the nature and properties of coke
11. Describe the nature and uses of lamp-black.
12. Illustrate the indestructibility of carbon.
13. What is meant by the absorbent power of charcoal?

14. Whence do the preservative properties of charcoal arise? Give illustrations.
15. Illustrate the decolorizing power of charcoal.
16. How does charcoal act as a reducing or deoxidizing agent
17. In what manner does charcoal exercise an oxidizing power? Give examples.
18. How may charcoal, impregnated with offensive odors, &c., be made useful?
19. What is meant by the term *allotropic?*

CHAPTER XI.

COMPOUNDS OF CARBON.

1. CARBON unites with oxygen in two proportions, forming two inorganic compounds of great interest, viz. :—

Carbonic Oxide	CO	=	28
Carbonic Anhydride	CO_2	=	44
Carbonic Acid	H_2CO_3	=	62

2. **Carbonic Acid Gas** (CO_2) is a colorless, inodorous gas, soluble in its own bulk of water, and about one and a half times as heavy as common air. It has a slightly acid taste, and reddens vegetable blues, but the blue color returns as the acid volatilizes. It extinguishes burning bodies, and is fatal to animal life.

3. Carbonic acid exists abundantly in nature. In the gaseous state it forms $\frac{1}{2500}$th part of the atmosphere; it is in solution in most spring water, and in some mineral waters to a considerable degree; it is found in the largest quantity, however, in solid combination with lime, magnesia, &c., for it is a constituent of all chalk, limestone, marble, coral, and shells.

4. Carbon and oxygen, as already stated, do not appear to combine by direct action at common temperatures: but compounds which contain carbon, when undergoing the

process of fermentation, yield a large quantity of carbonic acid. It is given off by plants during darkness; and is a constant product of respiration, of animal and vegetable putrefaction, and of combustion.

5. Carbonic acid may be formed (synthetically) by the combustion of any substance containing carbon; the several bodies named in the preceding chapter are shewn to be carbon from their each producing carbonic acid when burned.

Ex. a.—Burn charcoal in oxygen gas, as directed on page 27, carbonic acid will be generated.

Ex. b.—Hang a lighted taper in a bottle filled with common air, so long as it will burn. The carbon of the candle will have united with the oxygen of the air to form carbonic acid.

Fig. 31.

6. The presence of free carbonic acid may be tested by lime-water, with which it combines to form an insoluble carbonate of calcium or chalk, $Ca\ CO_3$.

Ex. c.—On removing the candle from the jar in the last experiment, pour in a little lime-water, and having closed the top of the jar,* shake it well, the transparent lime-water will soon become milky, from the formation of carbonate of calcium.

Ex. d.—Place a spoonful or two of fresh lime into a pint of rain-water, and let it stand for some hours, occasionally shaking it: the water will dissolve some of the lime and form the lime-water alluded to in the last experiment. After it has settled, pour off the clear liquid into a stoppered bottle, and keep it for future use.

Breathe through a tube into a wine-glass of lime-water, and the characteristic milky appearance will prove the presence of carbonic acid.

Ex. e.—Mix six grains of charcoal with 216 grains of red oxide of

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* If using a stoppered bottle, be careful not to close it with the stopper, otherwise, on the absorption of the gas by the lime-water, the atmospheric pressure will fix the stopper, and render its removal difficult.

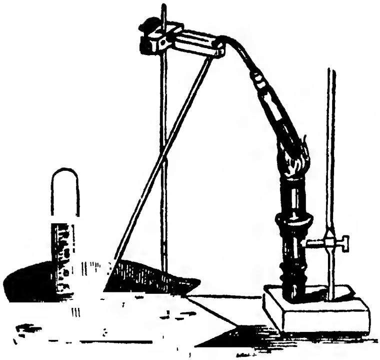


Fig. 32.

mercury, and heat them in a test tube, receiving the gas as in the diagram. This gas may be proved, by the taper and lime-water tests, to be carbonic acid. Had the oxide of mercury been used alone, oxygen would have been given off, as on page 24; under the present arrangement, however, instead of escaping, it combines with the carbon, forming carbonic acid, leaving the metallic mercury behind.

$$2HgO+C=2Hg+CO_2.$$

7. Carbonic acid gas is generally obtained (analytically) through the decomposition of a carbonate by the addition of an acid.

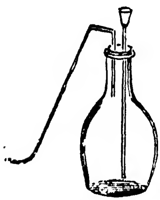


Fig 33

Fragments of chalk or marble (carbonate of calcium) are to be placed in such a bottle as was used in preparing hydrogen; and hydrochloric acid diluted with water is then to be poured slowly down the funnel. Carbonic acid gas will be rapidly evolved, and may be collected as usual.

$$CaCO_3+2HCl = CaCl_2+H_2O+CO_2.$$

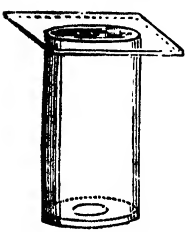


Fig. 34.

Ex. f.—Pour a little hydrochloric acid upon some pieces of chalk in a large-mouthed jar, effervescence arises from the escape of carbonic acid, which, on account of its density, will remain for some time in the jar: plunge a lighted taper into the gas, it will be immediately extinguished.

Ex. g.—Suspend a lighted taper in a jar, or, what is more striking, place several pieces at different depths, and then, having previously filled another jar A with carbonic acid, either at the pneumatic trough, or

by displacement, pour the gas from it as in the diagram, just as you would pour water, the tapers will be successively extinguished as the carbonic acid rises in the jar.

The experiment may be varied by ladling the carbonic acid from the jar in which it was prepared in the preceding experiment, into that in which the tapers are, with a small glass or tumbler for a ladle. Also by placing the delivery tube of the generating apparatus, fig. 33, so as nearly to reach the bottom of the jar B.

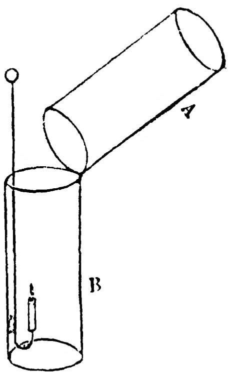


Fig. 35.

8. Owing to its great density or weight compared with common air, carbonic acid may be collected by displacement, as just shewn, that is by occupying the space previously filled by air, in the same way that water does when poured into an apparently empty vessel. For the same reason, it diffuses but slowly in the atmosphere, and is hence apt to remain a long time in old wells, brewers' vats, cellars, &c., and has frequently proved fatal to persons who have incautiously descended into such places. (*Choke-damp.*)

9. Fatal accidents occur not unfrequently from inhaling the fumes of charcoal, burned in close, ill-ventilated rooms. These fumes consist of carbonic acid, mingled with carbonic oxide, both gases being active poisons. Carbonic acid being a product of respiration, persons staying long in crowded apartments where there can be no current of air, will be destroyed by the accumulated gas, as in the well-known case of the Black Hole in Calcutta.

10. Carbonic acid does not, like nitrogen, cause death simply by excluding oxygen, and thus producing suffocation, but it is in itself a narcotic poison. Air containing a proportion of the gas insufficient to extinguish flame, will produce fatal effects if long breathed, gradually inducing stupor, insensibility, and death.

11. In order to resuscitate persons who have been exposed to the poisonous action of carbonic acid, cold water should be freely dashed upon them, their hands also and feet should be briskly rubbed, until proper assistance can be obtained.

12. In many places, where quantities of materials containing carbon are undergoing decomposition, streams of carbonic acid gas issue from the earth, particularly in the neighbourhood of volcanoes. There is a cave near Naples, the lower part of which is covered with a stratum of this gas, which proves fatal to dogs and small animals who are immersed in it, while men may walk upright without danger. The Poison Valley in the island of Java is a valley surrounded with mountains, and filled to a considerable depth with this deadly gas.

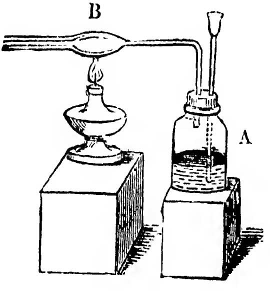


Fig. 36

13. Carbonic acid may readily be decomposed by heating a little potassium in an atmosphere of the gas. This may be done in a dry flask, filled with carbonic acid, by displacement; or as follows:—

Ex. h.—Adapt to the delivery tube of a generating bottle, A, a piece of a wider tube, or still better, one with a bulb, B, blown in it. Place some pieces of chalk in the gas bottle and pour over them some diluted hydrochloric acid; place also a globule of potassium in the bulb B. When the carbonic acid gas is passing over, apply a lamp, as in the diagram. The potassium will fuse and presently ignite, abstracting oxygen from the gas and leaving the black carbon plainly visible.

4

14. When subjected to a pressure of about 36 atmospheres $= 36 \times 15 = 540$lbs. per square inch, carbonic acid gas becomes a liquid lighter than water. The experiment is necessarily made in specially constructed vessels, gene-

rally of cast-iron. The liquefied gas, on being allowed to escape from the vessel in which it was formed, instantly returns to its gaseous condition, while a portion is immediately frozen by the cold produced by the rapid evaporation. A spirit thermometer placed in the jet of gas will sink to—90° F., and if a little ether be previously placed in the vessel, its vapor issuing with the gas will cause the temperature to sink to—194° F., the greatest degree of cold yet witnessed.

15. **Carbonic oxide,** CO, is a colorless, inodorous, tasteless gas, but slightly soluble in water, and fatal to animal life. It will not support combustion, but burns, in the presence of oxygen, with a blue lambent flame, carbonic acid being the product of the combustion.

16. Carbonic oxide may be obtained by passing carbonic acid through red-hot charcoal; the carbonic acid is decomposed, and yields half its oxygen to the burning carbon, $CO_2+C=2CO$.

17. This takes place in an ordinary fire. The coal at the bottom of the grate, where the air is plentiful, receives its full supply of oxygen, and becomes carbonic acid: the heated fuel above, where the quantity of air is limited, takes half the oxygen from this carbonic acid, and reduces it to the state of carbonic oxide, which passes to the top of the fire, and there, where air is once more abundant, burns with its characteristic blue flame,* being again converted into carbonic acid. The several steps may be shown, as follows:—

1st step, $C+2O= CO_2$, at the bottom of the grate
2nd step, $CO_2+ C=2CO$, passing through the fire.
3rd step, $2CO+2O=2CO_2$, on the top of the fire.

18. Carbonic oxide may be obtained by heating a carbonate such as chalk, with iron filings, or with charcoal;

* It may be readily observed in a charcoal stove, or when an ordinary fire is burning clear or frosty, as it is sometimes termed.

in either case the carbonate is decomposed, the iron or the charcoal absorbing a portion of the oxygen from the carbonic acid.

$$\text{Thus:—}CaCO_3 + Fe = CaO + FeO + CO.$$
$$\text{or}\quad CaCO_3 + C = CaO + 2CO.$$

19. The most convenient mode of preparing carbonic oxide is by *gradually* heating some oxalic acid, $H_2C_2O_4$, in a large retort, with about five or six times the weight of strong sulphuric acid. The oxalic acid is decomposed into carbonic acid, carbonic oxide and water; by passing the gases through lime-water in a Woulfe's bottle, the carbonic acid may be absorbed, leaving the carbonic oxide to be received at the trough in the usual way.

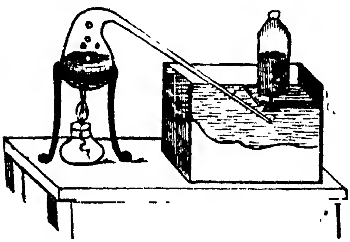


Fig. 37.

20. The reaction arises as follows:—Oxalic acid contains one atom of carbonic oxide, one atom of carbonic acid gas, and one atom of water: moreover the acid cannot exist except in combination with water or some base; the strong sulphuric acid deprives it of its water of crystallization, and the gases are immediately set free; thus—

21. Pure carbonic oxide may be obtained at once by the action of strong and hot sulphuric acid (H_2SO_4) on the ferrocyanide of potassium or prussiate of potass (K_4FeCy) or ($K_4FeC_6N_6$).

The decomposition is as follows :—

$$K_4FeC_6N_6+6(H_2SO_4+H_2O)=$$

$$2(K_2SO_4+3\{NH_4)_2SO_4\}+(FeSO_4)+6CO$$

That is :—

Ferrocyanide of potassium and strong sulphuric acid yield, 2 of sulphate of potassium, 3 of sulphate of ammonium, 1 of protosulphate of iron, and 6 of carbonic oxide.

Ex. i.—For small experiments, the gas may be readily prepared in a test tube, by covering a little oxalic acid with oil of vitriol, and heating them together; on the gas rising, it may be ignited at the end of the tube, as in fig. 30.

EXERCISES ON CHAPTER XI.

1. Name the principal compounds of oxygen and carbon.
2. State the properties of carbonic acid.
3. Name some natural combinations of carbonic acid.
4. How is carbonic acid generated in nature ?
5. Illustrate the synthetical formation of carbonic acid.
6. What is a test for the presence of free carbonic acid ?
7. Give some analytical process for obtaining carbonic acid.
8. What consequences arise from the great density of carbonic acid ?
9. What are the injurious effects of ill-ventilated rooms ?
10. How does carbonic acid differ from nitrogen in causing the death of those who inhale it ?
11. What means of restoration should be employed in the case of persons poisoned by carbonic acid gas ?
12. Where may quantities of carbonic acid gas be found free in Nature ?
13. How may carbonic acid be decomposed ?
14. Explain the liquefaction and solidification of carbonic acid gas.
15. State the properties of carbonic oxide.
16. How may carbonic oxide be obtained ?
17. Describe the process of combustion in a common fire.

18. Explain the mode of obtaining carbonic oxide from a mixture of chalk and iron filings.
19. Describe the process of obtaining carbonic oxide from oxalic acid.
20. Describe the reaction that takes place on the decomposition of oxalic acid.
21. How may carbonic oxide be obtained without any mixture of carbonic acid?

CHAPTER XII.

CARBON AND ITS COMPOUNDS—*continued.*

1. **Carbon** combines with Hydrogen to form a very numerous class of compounds, termed **Hydro-carbons,** which are all highly combustible. Most, if not all, of these belong strictly to organic chemistry, inasmuch as they always arise from the decomposition of organic bodies.

2. Two of these hydro-carbons, however, exist ready-formed in nature, viz.:—

Light Carburetted Hydrogen, (*marsh-gas,* or *fire-damp,*) $CH_4 = 16$.

Heavy Carburetted Hydrogen, (*olefiant-gas,*) $C_2H_4 = 28$.

3. **Light Carburetted Hydrogen** is a colorless, tasteless, and nearly inodorous gas, about half as heavy as common air, and but slightly soluble in water. It cannot support respiration, though it does not appear to be actually poisonous

4. Like pure hydrogen, the light carburetted hydrogen will not support combustion; but unlike that element, which may be inflamed by a red-hot wire, it cannot be

ignited till raised to a white heat, when it burns with a yellowish flame: the products of its combustion in common air, being water and carbonic acid; thus—

$$CH_4 + 4O = CO_2 + 2H_2O$$

5. Light carburetted hydrogen is formed abundantly in the mud of stagnant pools and marshes, by the decomposition of vegetable matter *under water*, hence it has been named *marsh gas*.

6. It may be obtained from this source by stirring up the mud at the bottom of a pond, and collecting the bubbles of gas, as they rise, in a jar or bottle filled with water and inverted over them. The gas thus procured is almost always mixed with a little carbonic acid and nitrogen, arising from the decaying vegetable matter. The carbonic acid can be removed by agitating the gas in a bottle with lime-water, while the small quantity of nitrogen will not interfere with the exhibition of the peculiar properties of the marsh-gas.

7. As coal is without doubt a substance of vegetable origin, it is not surprising that light carburetted hydrogen is frequently found in coal mines, issuing from fisssures in the strata or seams of coal, in consequence of the decomposition of that mineral.

8. When mingled with from six to fourteen times its bulk of air, it forms the dangerously explosive mixture, termed by the miners, *Fire-damp*.

Carbonic acid, as already stated, is produced by the combustion, so that those miners who escape the immediate effects of the fire or the shock, are often suffocated by this poisonous and deadly gas, known by the miners as the *after-blast* or *choke-damp*.

9. Light carburetted hydrogen may be obtained in purity by strongly heating in a retort or flask a mixture of

2 parts of crystallized acetate of sodium ($NaC_2H_3O_2 + 6HO$), 2 parts of solid hydrate of potassium or caustic potass (KHO), and 3 parts of quicklime in powder, CaO.

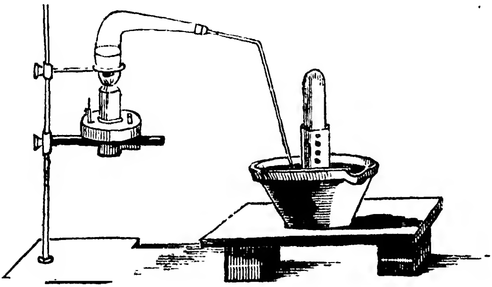


Fig. 38.

The gas, which is disengaged in abundance, may be collected over water, as usual.

10. The acetate of sodium is first rendered *anhydrous* by the heat, and then decomposed, the elements of each atom of acetic acid uniting with the oxygen and hydrogen of one molecule of water and rearranging themselves into two atoms of carbonic acid (which remain in the retort combined with the potassium and sodium), and two atoms of light carburetted hydrogen. The use of the lime would seem to be to prevent the hydrate of potassium from attacking and fusing the glass vessels.

The reaction may be expressed in the following formula:

$$2(NaC_2H_3O_2) + 2KHO = Na_2CO_3 + K_2CO_3 + 2CH_4.$$

Ex. a.—Fill a capped glass jar with the gas, screw a jet into the stop-cock, and depress the jar into the water of the pneumatic trough, so as to force a stream of gas through the jet. Endeavour to inflame the gas by a red-hot wire, it will be found impossible to do so; apply a lighted taper, and the gas will burn with a pale flame, owing to the comparatively little carbon it contains (see page 34, § 16).

Ex. b.—Perform a similar experiment with hydrogen and carbonic

oxide, and observe that each of these gases inflames readily at a red heat.

Ex. c.—An impure light carburetted hydrogen is obtained from wood by simple heating. For this purpose sawdust or bits of shavings may be heated in a test-tube, and the gas burned in a jet, as fast as it is formed; fig. 30. The flame is more luminous on account of the gas being mixed with heavy carburetted hydrogen, &c.

11. **Heavy Carburetted Hydrogen, C_2H_4,** is frequently termed **Olefiant* gas,** from its property of forming an oily liquid when mixed with an equal volume of chlorine, even in the dark. The substance so formed is termed *Dutch liquid,* from having been discovered by some Dutch chemists; it is a heavy, volatile, colorless liquid, possessing a sweetish taste and an agreeable odor.

12. Olefiant gas is prepared by mixing strong alcohol or spirits of wine in a large flask or retort, with five or six times its weight of sulphuric acid, and applying heat. At first alcohol, and then ether, distil over; but as the temperature rises, the mixture darkens, and gaseous products are evolved in abundance. These are olefiant gas, carbonic acid, and sulphurous acid; the two last-named substances may be removed by passing the mixed gases through lime-water, as they come from the retort, or by allowing them to stand over water for some time.

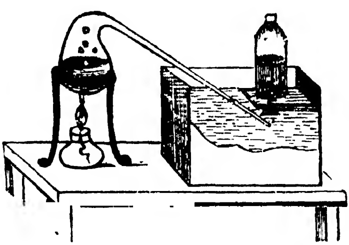


Fig. 39.

* Olefiant—*olefacient,* oil-making, from *oleum,* oil, and *facio,* I make.

13. A mixture of sulphuric acid and alcohol forms a peculiar compound (Sulphovinic acid, $C_2H_6SO_4 + H_2O$), which is readily decomposed by heat, yielding different products, according to the temperature employed.

If the mixture boil at a temperature—

Below 260°, the alcohol distils over unchanged.

Between 260° and 310°, ether and hydrated sulphuric acid are formed.

Above 320°, olefiant gas, &c. are produced as above mentioned.

14. These changes may be explained by the affinity subsisting between sulphuric acid and water. When alcohol is heated with excess of sulphuric acid, so as to raise the boiling point above 320°, the acid separates the water from the olefiant gas, and the latter is given off.

$$C_2H_6O + H_2SO_4 = (H_2SO_4 + H_2O) + C_2H_4.$$

The blackening of the mixture results from a secondary action of the sulphuric acid upon a portion of the alcohol, by which carbonic and sulphurous acids with other products are formed while some of the carbon is deposited.

15. Olefiant gas is colorless and tasteless, soluble in about 8 times its bulk of water, and somewhat lighter than atmospheric air. It is fatal to animal life, and very combustible, burning with an intensely bright flame. When exposed under pressure to a degree of cold 166° below 0°, it can be condensed into a transparent, colorless liquid.

16. For complete combustion:—

One volume of marsh gas requires 10 volumes of air, or 2 volumes of oxygen.

One volume of olefiant gas requires 15 volumes of air, or 3 volumes of oxygen.

17. The light emitted by ordinary coal-gas is due to the heavy and light-carburetted hydrogen which it contains: its illuminating power being in proportion to the quantity of the former. Various volatile hydro-carbons increase the luminosity of the flame, and give to the gas a peculiar odor.

18. Gas for illuminating purposes is generally prepared from bituminous coal, by subjecting it to destructive distillation in cast-iron retorts maintained at a bright red heat.

19. A variety of gaseous and vaporous products are given off during the process, the chief of which are olefiant gas, light carburetted hydrogen, sulphuretted hydrogen, carbonic acid, ammonia, and tar, while the carbon of the coal remains in the retort in the form of coke.

20. The amount and character of products just mentioned vary in proportion to the temperature to which the coal, &c., is exposed. The gases become more abundant with an increase of heat, but less valuable, on account of their losing carbon, and therefore burning with a less luminous flame.

21. Coal gas, as it issues from the retorts, cannot be directly employed for illumination, since it contains products highly offensive to the smell, as well as some that would condense and block up the pipes through which they might have to pass.

22. The gas, &c., on leaving the retort, is received in a large horizontal pipe half filled with water (*the hydraulic main*); from this it is made to traverse a series of iron pipes kept constantly cool (*the refrigerator*), till it reaches a vessel termed *the condenser;* during this course the ammonia, tar, and other vapors have been condensed.

23. From the condenser the gas passes onward into another chamber (*the purifier*), where it is forced through

a cream-like mixture of lime and water, which removes the sulphuretted hydrogen and carbonic acids.

The gas, thus purified, is conducted into large reservoirs (*gasometers*), whence it is conveyed by pipes to the spots where it may be required.

24. A gas (*oil-gas*) richer in heavy carburetted hydrogen, and therefore more valuable for illuminating purposes, may be obtained simply by dropping refuse oil, or resin, into red-hot retorts filled with coke. The gas so prepared is free from the sulphurous and ammoniacal compounds contained in coal gas, but the expense of the materials has prevented its general adoption.

25. According to Dr. Ure's statement,

1 pound	of good	coal yields	4	cubic	feet	of illuminating	gas.
1	,,	,, resin ,,	10	,,	,,	,,	,,
1	,,	,, oil or fat	15	,,	,,	,,	,,

26. The same author computes that an amount of light produced from wax for one shilling, would cost 3½d. from tallow, 1¾d. from oil, and somewhat more than ½d. from coal-gas.

"One pound of mould candles (six to the pound), if burned in succession, will last 40 hours; 11½ cubic feet of gas, burned at 500 cubic inches per hour, will give the same light for the same time."

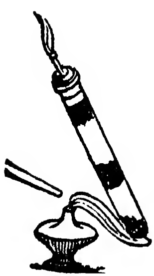


Fig. 40.

Ex. d.—Place a little coal at the bottom of a long test tube, and above it a loosely fitting plug of rag or paper moistened with water; still nearer the mouth of the tube place another plug moistened with sugar of lead. Heat the coal by a lamp or blow-pipe flame, it will be decomposed; the first plug containing water will retain the ammonia, the second will retain the sulphuretted hydrogen, and become black from the sulphuret of lead formed. The gas as it issues from the jet may be readily inflamed.

Ex. d.—Fill the bowl of a common tobacco-pipe with coal-dust, cover it with sand or clay, and place it in the fire; when hot, carbu-

retted hydrogen gas will be evolved, and may be lighted at the end the stem of the pipe.

Ex. f.—Mix together 3 parts of oxygen and 1 of coal-gas Blow soap bubbles with the mixture, (as in *ex. f.* page 34), and inflame them by a candle; they will detonate with a very loud report.

Ex. g.—Mix together in a stout soda-water bottle 1 volume of coal gas and 10 of atmospheric air, cork the bottle and wrap it in a cloth to prevent accidents. Then uncork the bottle, and holding the mouth away from you, apply a flame to the mixed gases, explosion wll take place. This shows the care requisite in entering any apartment, in which an escape of gas has taken place, with a lighted candle.

Ex. h.—Mix together in a tall glass jar 1 volume of olefiant or coal-gas, and 2 of chlorine; inflame them by a lighted candle, the result will be hydrochloric acid, with a deposition of charcoal. In this experiment the coal-gas is decomposed, its carbon falls down, while its hydrogen, uniting with the chlorine, forms the acid mentioned (chap. xvi).

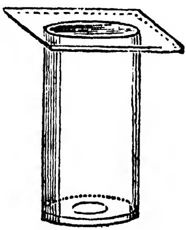


Fig. 41.

Ex. i.—Let a jar half full of olefiant gas be placed on the shelf of the pneumatic trough, and then filled up with chlorine gas. The gases will gradually unite, forming an oil-like substance, which floats on the water as it ascends in the jar. This shews the cause of the name *olefiant.*

EXERCISES ON CHAPTER XII.

1. What are hydro-carbons?
2. Name the most important hydro-carbons.
3. State the properties of light carburetted hydrogen?
4. What temperature is required for its ignition? What are the products of its combustion in common air?
5. Why has it been termed *marsh-gas?*
6. How may it be naturally obtained?
7. In connection with what mineral is it found?
8. What is meant by *fire-damp* and *choke-damp?*
9. How may *marsh-gas* be artificially prepared?
10. Describe the reaction that takes place in the above process.
11. What are olefiant gas, and Dutch liquid? Whence do they derive their names?
12. How may olefiant gas be prepared?

13. State the peculiarities in the decomposition of a mixture of alcohol and sulphuric acid.

14. How are the changes observed in the production of olefiant gas to be accounted for?

15. State the properties of olefiant gas.

16. State the quantities of air or oxygen necessary for the complete combustion of marsh gas and olefiant gas.

17. To what is the illuminating power of coal-gas due?

18. How is illuminating gas prepared?

19. What are the chief products of the destructive distillation of coal?

20. What effect has the heat of production on the luminous qualities of coal-gas?

21. Why does coal-gas need purification?

22. How are the vaporous compounds in coal-gas removed?

23. Why is coal-gas passed through lime-water?

24. How is oil-gas prepared, and in what respect does it differ from coal-gas?

25. State the comparative quantity of gas given from coal, resin, and oil.

26. State the comparative cost of light produced from the combustion of different bodies.

CHAPTER XIII.

FLAME.

1. **Flame** is the combustion of gas or vapor raised to so high a temperature as to emit light as well as heat.

2. Two conditions are necessary for illuminating flame, viz.:—a sufficiently high temperature, and the presence of solid matter within the heated space.

3. The temperature may be very high, when the light is feeble, as is seen in burning together pure oxygen and hydrogen, the flame being scarcely visible in the day-time, though its heat is intense. (See Chapter V. §§ 15, 16.)

4. The elements of wood, tallow, oil, illuminating gas, &c., with which oxygen unites in ordinary combustion, are chiefly carbon and hydrogen; the former of these escapes

into the air as carbonic acid, the latter as water in the state of vapor.

5. Ordinary flame is not a solid mass of fire, but rather a *hollow*, luminous cone, in the centre of which there is no combustion.

6. In the burning of a candle, (which affords one of the best flames for illustration,) the wax or tallow being first melted by the heat, rises in the wick by the force of capillary attraction: and although the wick supplies some oxygen and hydrogen, by far the greater part of these is furnished by the decomposition of the wax or tallow, which burn in connexion with the oxygen of the air. The supply of heated vapor diminishes as it ascends, and eventually ceases, and hence the flame of a candle tapers to a point or becomes conical.

7. A flame, if carefully observed, will be found to consist of three distinct parts, viz.:—

A dark central space, *b*; a *highly* luminous cone enveloping it; and on the exterior a second *faintly* luminous cone, *aa*, having a high temperature.

8. The dark central space, in which there is no combustion, is filled with unburned gas arising from the wick.

Ex. a.—This may be proved by inserting one end of a small glass tube into the dark portion of the flame of a large candle; the unburnt vapor will escape through it, and may be ignited at the other end of the tube.

Ex. b.—The fact that the flame is hollow may also be rendered apparent by stretching across it a fine iron or platinum wire, when the central portion of the wire will appear dark, while the parts near the exterior of the flame are raised

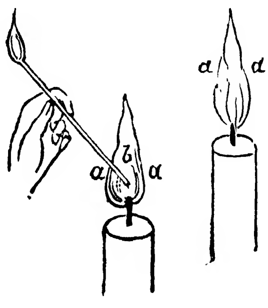


Fig. 42

to a white heat. In this way we can readily ascertain where the temperature of the flame is greatest.


Fig. 43.

Ex. c.—Hold a piece of wire gauze over the flame, and while combustion is going on only below, look down upon it and observe that in the centre there appears a dark spot, from which gas rises through the gauze, surrounded by a ring of light.

9. The highly luminous cone surrounding the dark centre consists of carburetted hydrogen, or illuminating gas, the hydrogen of which burns *first*, while the great heat evolved raises the particles of carbon to a state of intense ignition.

10. In the faintly luminous exterior cone, the carbon particles are consumed, forming carbonic acid with the oxygen of the air.

Ex. d.—That the luminous part of the flame consists of intensely ignited particles of charcoal, may be shewn by introducing into it any cold body, as a knife, a glass rod, an earthenware plate, or an iron spoon, which will become blackened by the decomposition of the lamp-black or carbon.

11. The reason the carbon particles are not consumed *with* the hydrogen in the highly luminous cone may be thus explained. Carbon and hydrogen differ in their attraction for oxygen, the latter greatly exceeding the former; consequently, when both are present, and the supply of oxygen is limited, the hydrogen takes all, leaving the carbon to burn where it can obtain a further supply of oxygen, which it does in the exterior flame.

12. If oxygen be mingled with, or driven into the flame as in the oxy-hydrogen light (Chap. V. §§ 18, 19), or by the ordinary blowpipe, the hydrogen and carbon burn *together*, and the luminosity of the flame almost entirely disappears.

13. In an Argand burner, whether oil or gas, air is permitted to enter through the centre of a circular wick, or a

ring of gas jets, so that the intensity of heat is increased, more gas is consumed in the same space than by any other mode of combustion, more carbon therefore is precipitated, and this being more intensely heated, greatly increases the luminosity of the flame.

14. If the central opening through which the air is admitted be closed, part of the supply of oxygen necessary for complete combustion is withdrawn; the flame becomes elongated as if in search of air, the heat is less intense, less carbon is consumed, and being at a lower temperature, with a feebler light; the unconsumed carbon gives a yellowish tinge to the flame, and is finally deposited in the form of lamp-black: in short, the lamp *smokes*.

15. The effect of diminished temperature on the brilliancy, and on the very existence of flame, may be further illustrated as follows:—

Ex. f.—Form a small coil of metal, by winding some copper wire several times round a pencil; place the coil gently over the wick of a wax or composite candle, so as not to touch it, the flame will be extinguished. The mass of metal being a good conductor of heat, abstracts it so rapidly from the burning gas, that the latter can no longer maintain itself at a white heat, the temperature necessary for flame. If the coil be *heated* before trying the experiment, no such effect will follow.

16. A piece of wire gauze placed over a flame, as in *Ex. c.*, arrests it, not by stopping the passage of the burning gas, but by cooling it below the necessary temperature, as it passes through.

Ex. g.—Intercept a flame by a piece of wire gauze, and apply a light to the upper surface, the unconsumed gases which were rising through the meshes of wire will be inflamed.

Ex. h.—Hold a piece of wire gauze over a jet from which gas is issuing, apply a light above the gauze, the gas as it passes through will burn, but will not inflame that beneath. By gently raising the gauze the flame may be removed several inches above the gas jet.

17. The foregoing experiments illustrate the principle

of the safety lamps invented by George Stephenson, the celebrated engineer, and by Sir Humphrey Davy, with a view to prevent the fatal explosions of fire-damp so frequent in coal-pits. The *Davy* lamp consists of an ordinary oil-lamp, *b*, surrounded by a cage or casing of wire gauze. Oil is supplied through the opening at *a*, and the wick can be regulated by the wire *c*, without opening the gauze casing. Fire-damp can only be inflamed when raised to a white heat; if it come into contact with an unshielded candle, it of course inflames and explodes immediately. The safety-lamp, however, guards against this explosion by confining the combustion of the mixed gases forming the fire-damp within the gauze cage, and preventing the flame from communicating with the rest of the fire-damp outside.

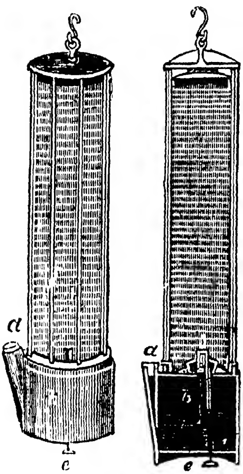


Fig. 44.

18. The safety-lamp must be regarded not only as a security against sudden explosion, but also as conveying a warning of danger, and affording time for escape; for if the wire gauze should, by the ignition of a highly combustible mixture within, become ignited to whiteness, the exterior gas would of course take fire, and explosion follow.

19. Hemming's jet for burning the mixed gases, already referred to (Chap. V. § 31), acts on a similar principle to that of the safety-lamp, the cooling power of the meta tubes formed by the wires being so great as to preven the passage of flame

20. The inner luminous cone (§ 9), with the heated gas within, has *deoxidizing* effects—that is, it will *reduce* metallic oxides exposed to it to the metallic form. This it does because it contains much carbon and hydrogen, which take the oxygen from the oxide, passing off as carbonic acid and water, and leaving the metal behind.

21. The exterior faintly luminous cone, on the contrary, has an *oxidizing* effect, owing to the abundant supply of heated air near it.

Ex. i.—Hold a strip of tarnished copper upright in the flame of a spirit-lamp or candle, the part within the flame will lose its tarnished appearance, while the outer portion will become more deeply oxidized, and consequently darker. These changes may be readily observed by moving the metal to and fro, and noticing the variety of color produced.

22. The peculiar effects of the oxidizing and deoxidizing flames may be still better obtained by means of a blow-pipe or tube, through which air can be conveyed into the central space of unconsumed gases. (§ 8.)

23. These gases being thus supplied with air, combustion goes on both in the centre of the flame and at its outer border, the heat is necessarily increased, and the flame elongated, consisting of two cones—the inner, blue; the outer, yellow.

24. Between these two cones there is still a portion of the flame where consumed gas is present, as in § 9, and where the amount of white-hot carbon is ready to abstract oxygen from any oxide exposed to it. This part of the flame is therefore termed the *reducing* flame, and anytning to be reduced must be held *within* it.

25. The extremity of the outer or yellow cone of the blow-pipe flame is termed the *oxidizing* flame, having a similar effect to the faintly luminous envelope of an ordinary flame, § 10. Substances to be oxidised should be held a little in front of the visible apex, for it is the heated air rather than the actual flame which produces the oxidation.

EXERCISES ON CHAPTER XIII.

1. Give a definition of flame.
2. What are the conditions necessary for illumination ?
3. Illustrate the relations between the light and heat of flame.
4. What becomes of the elements of a burning candle ?
5. What is the character of ordinary flame ?
6. Trace the progress of combustion in a candle. Give a reason for the conical form of its flame.
7. Into what parts may an ordinary flame be divided ?
8. In what part of the flame is there no combustion ? Give proofs.
9. Account for the luminosity of the inner cone.
10. Where are the carbon particles consumed ?
11. Why is the carbon not consumed in the inner cone ?
12. What would be the effect if the carbon and hydrogen burned *together ?*
13. What is the principle of an Argand burner ?
14. Explain the effect of diminished draught on an Argand burner.
15. Illustrate the effect of temperature on flame.
16. Why does wire gauze intercept the passage of flame ?
17. Describe the safety-lamp.
18. When is a safety-lamp unsafe ?
19. What is the principle of Hemming's oxy-hydrogen jet ?
20. Why does the body of an ordinary flame *reduce* metallic oxides ?
21. How is the oxidizing effect of the exterior envelope of an ordinary flame accounted for ?
22. How may the chemical effects of flame be more readily applied ?
23. What is the effect of blowing air into the centre of a flame ?
24. What is meant by the reducing flame of the blow-pipe ?
25. Where is the greatest oxidizing power of the blow-pipe flame ?

CHAPTER XIV.

SULPHUR AND SELENIUM.

Sulphur. *Symbol* S; *combining weight* 32
Selenium. ,, Se; ,, ,, 79.5

1. **Sulphur**, or *brimstone*, is a yellow, brittle, crystalline, and combustible, solid; tasteless, inodorous (except when rubbed), and perfectly insoluble in water.

2. Sulphur is found pure or *native* in different parts of the world, often in connexion with beds of gypsum and rock-salt, but more especially in volcanic districts. The greater part of the sulphur of commerce is obtained from the island of Sicily, where it is quarried from large deposits.

3. Sulphur in a state of combination is universally diffused throughout nature. It is found in plants and animals, while in the mineral kingdom it forms the numerous and important compounds termed *sulphurets* or *sulphides*, and *sulphates*.

4. Sulphur melts readily at a temperature of 226° F., forming a pale yellow liquid; at 300°, it begins to thicken and turn brown; at 450°, it changes to a thick, tenacious, treacle-like body; at 480°, it becomes once more a thin liquid; and at 600°, it boils, giving off a deep reddish-brown vapor, which is condensed by cold into a fine yellow powder, termed *flowers* or rather *flour of sulphur*.

5. If the thick, tenacious, treacle-like substance just mentioned be poured into water, it becomes soft and elastic, like indian-rubber, and may be used for taking impressions of coins, seals, &c. This is an *allotropic* form of sulphur; after a while it becomes once more brittle and crystalline.

6. The process by which a solid substance is evaporated and again condensed into the solid form, is termed *sublimation*, as distinct from *distillation*, in which the vapor is condensed into the form of a liquid.

Ex. a.—Melt a quantity of sulphur in a crucible or deep iron ladle, then set it aside to cool; when the surface is congealed, pierce the crust with an iron wire, and pour out the still liquid sulphur beneath; let the crust, &c. remain till perfectly cold, and then, on turning it out of the crucible, the sulphur will be found in long needle-shaped crystals.

Ex. b.—The union of sulphur with a metal, and thereby forming a sulphuret, has been already noticed (*Chap. III. Ex. e.*)

7. The purity of sulphur may be judged of by heating a portion on a piece of platinum foil; if pure, the sulphur will burn away without leaving any residue. Common *flowers of sulphur* are sufficiently pure for all ordinary purposes. *Roll sulphur* is merely sulphur melted and cast in cylindrical wooden moulds.

8. Oxygen and sulphur form several compounds, all of which possess acid properties, for there is no neutral oxide of sulphur known. The principal of these compounds are as follow:—

	Anhydride.		Acid.
Sulphurous . .	SO_2	;	H_2SO_3.
Sulphuric . .	SO_3	;	H_2SO_4.
Hyposulphurous.	S_2O	;	$H_2S_2O_3$.
Hyposulphuric .	S_2O_5	;	]

9. **Sulphurous acid gas**, SO_2.—When sulphur is heated in the air or in pure oxygen to about 300°F., it inflames, and combining with the oxygen, forms a colorless, transparent gas, having a peculiarly disagreeable taste and a most suffocating odor: this gas is sulphurous acid, and is given off every time a sulphur match is ignited.

10. Sulphurous acid gas is more than twice as heavy as common air; it extinguishes burning bodies, is uninflam-

mable, and instantly fatal to animal life.* Cold water dissolves about 35 times its bulk of the gas, which is evolved unaltered when the solution is heated. The aqueous solution, like the gas itself, has the property of bleaching some animal and vegetable substances. Hence the vapor of burning sulphur is used for whitening, silk, woollen or straw goods, discolored paper, and numerous other articles.

Fig. 46.

Ex. c.—Burn sulphur in oxygen (as shown *Chap. IV. Ex. c.*), observe the formation of sulphurous acid; when the sulphur has burned out, introduce a lighted taper into the jar, the flame will be instantly extinguished.

Ex. d.—Immerse a piece of moistened litmus paper in a jar of the gas, or hold it in the fumes of burning sulphur, it will at first be reddened and afterwards slowly bleached.

Fig. 47.

Ex. e.—Dip a bunch of violets, a red rose, peony, dahlia, or some other flower, into a solution of sulphurous acid, or suspend it in the fumes of the gas, as in the diagram: the color will be speedily discharged. Moisten the bleached flower with some ammonia, the action of the sulphurous acid will be neutralized, and a green color will be produced.

Ex. f.—Pass some sulphurous acid through an infusion of red cabbage, which has been rendered green by mixing with it a little soda or potass; the solution will gradually redden, and at length become colorless. Divide the colorless liquid into two portions, add a little sulphu-

* "For this reason burning sulphur is used to destroy bees in the hive previous to taking the honey. A bundle of lighted matches will effectually destroy all the bees of a hive, sometimes 20,000 in number. Insects to be preserved for thecabinet are mostly killed by the fumes of sulphur, though the fumes of prussic acid are preferable, because sulphur sometimes injures their colors."—*Francis.*

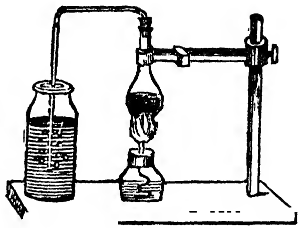


Fig. 48.

ric acid to one, a red color will be produced; add some potass or soda to the other, and it will become green. This shews that the coloring matter, though *rendered invisible* by the gas, was *not destroyed.*

11. Sulphurous acid gas may be readily liquified at the ordinary pressure of the atmosphere, by passing it through tubes surrounded by a mixture of salt and snow, which produces a temperature of 0°; or by subjecting it at the freezing point of water, 32°, to the pressure of two atmospheres.

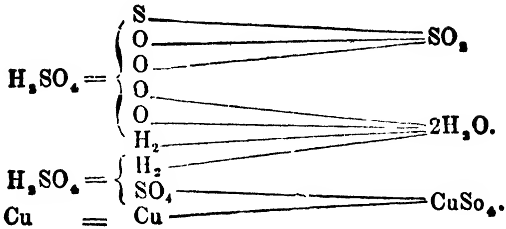
12. Sulphurous acid gas may be readily obtained by heating copper turnings, pieces of copper wire, or metallic mercury in a retort or tube with sulphuric acid. A portion of the acid is decomposed, and gives up one third of its oxygen to oxidize the copper; this is immediately dissolved by another portion of acid, and forms sulphate of copper; the sulphur of the decomposed sulphuric acid is set free in combination with the remaining two-thirds of oxygen, forming sulphurous acid gas, which may be collected over warm water, or by *displacement.*

(*Compare the process for obtaining Binoxide of Nitrogen, Chap. VIII.* §§ 13, 14.)

13. Two atoms of acid are necessary to each atom of copper, to produce a perfect decomposition, thus—

$$2(H_2SO_4)+Cu = CuSO_4+2H_2O+SO_2.$$

or, more fully, as follows:—



14. Aqueous sulphurous acid may be readily formed by passing a current of gas through cold water, as shown in fig. 48. This solution may be kept unchanged as long as air is excluded; but when exposed to oxygen it slowly combines with it, and forms sulphuric acid. The compounds of sulphur*ous* acid are termed sulph*ites*.

15. **Sulphuric acid, H_2SO_4.** This powerful acid has been long known and extensively employed in its *hydrated* state, that is in chemical combination with water. Two varieties of the acid are known in commerce, both of which are liquid. The *anhydrous* acid, as it is termed, is merely an object of curiosity, and does not appear to possess acid properties (*see Chap. VII.* § 15). The names, symbols and equivalents of these three compounds, are as follows :—

Anhydrous sulphuric acid	SO_3	= 80.
Hydrated sulphuric acid (*oil of vitriol*)	H_2O,SO_3	= 98.
Nordhausen or fuming sulphuric acid,	$H_2O,2SO_3$	= 178.

16. Sulphuric acid was formerly obtained by the distillation of sulphate of iron, $FeSO_4 + 7HO$, called by the alchemists, *green vitriol ;* hence the acid is still sometimes termed *vitriolic* acid, though from its oily appearance, it is most commonly known as *oil of vitriol.*

17. This process is still followed at Nordhausen in Saxony, hence the acid produced by it is termed Nordhausen acid ; its other name of *fuming* sulphuric acid arises from

the fact of its emitting white vapors of hydrated sulphuric acid, H_2SO_4, by uniting with the moisture always present in the atmosphere.

The Nordhausen acid is chiefly prepared for the purpose of dissolving indigo, the solution so formed being very extensively used as a blue dye.

18. In order to procure the Nordhausen acid, the sulphate of iron, $FeSO + 7H_2O$, is strongly heated, so as to expel the greater part of its water of crystallization. The nearly anhydrous sulphate is then heated to redness in earthenware retorts, by which means it is decomposed; half its sulphuric acid is resolved into sulphurous acid gas and oxygen; the latter unites with the protoxide of iron, to form the sesquioxide; while the other half of the sulphuric acid comes over in combination with water, thus—

$$4FeSO_4 + H_2O = 2Fe_2O_3 + 2SO_2 + H_2O,2SO_3.$$

19. **Sulphuric Anhydride, SO_3.** — This substance may be procured by distilling strong Nordhausen acid at a low temperature, and receiving the vapor in a receiver kept very cool. It condenses into white, fibrous, solid, silk-like crystals, which liquefy at 66°, and boil at about 86°, forming a transparent vapor, provided no water is present.

20. **Hydrated sulphuric acid, H_2SO_4.** — Sulphuric acid in its most concentrated form is a hydrated acid, containing one equivalent of water, and is, perhaps, the most important of the acids. It cannot be procured by the direct combination of its elements, like carbonic or phosphoric acids, which may be obtained by the simple combustion of carbon or phosphorus in oxygen. The product of the combustion of sulphur in oxygen is, as we have already seen, sulphurous acid gas; this, however, may be oxidized by mixture with nitric acid vapor, and so converted into sulphuric acid.

21. Sulphuric acid is prepared on a large scale by heating sulphur and nitre in furnaces, and conducting the sulphurous acid and nitric acid fumes along with steam and atmospheric air into leaden chambers, the floors of which are covered with water.

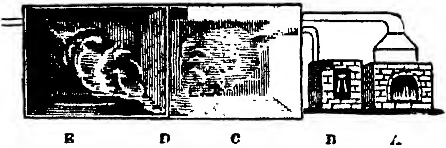


Fig. 49.

A is a boiler, to afford steam. B is a small furnace in which sulphur is ignited. Over the sulphur is another vessel, supported on a stand, and containing sulphuric acid and nitrate of potass, the materials for making nitric acid, the formation of which commences as soon as the burning sulphur has imparted the requisite heat. The sulphurous and nitric acid vapors thus produced pass with the steam from the boiler A into the leaden chamber, which is divided into two parts, C and E, by a partition D, nearly reaching to the floor. F is a pipe to carry off any uncondensed gases, and by means of it a slight current of air is maintained through the chamber, causing the gases to be more thoroughly intermingled as they pass below the partition (or alternately above and below, in a series of chambers). In this way the water on the floor soon becomes acid; it is then drawn off, heated to 300° F. in leaden troughs, to evaporate part of the water, and finally concentrated in platinum stills, till it acquires a specific gravity of about 1.850 (water being 1), and boils only at 618°. Lead is used as a lining to the chambers, &c. because almost any other material would be destroyed by the acid, and even the lead itself is corroded by it when strong; hence, notwithstanding the expense, platinum stills, costing from £1000 to £2000, are used to concentrate the acid, being preferred to glass retorts, on account of their durability.

Ex. g.—Sometimes, instead of forming the necessary nitric acid at the moment, vessels containing that acid are placed in the first chamber in the current of sulphurous vapor. Ignite some pieces of sulphur in a deflagrating spoon, and suspend them in a bottle containing a little water. After the flame has died out, hold in the sulphurous acid gas which fills the bottle a piece of flannel fastened to a glass rod and moistened with nitric acid, orange-coloured fumes—hyponitric anhydride—will immediately form, the sulphurous acid will be gradually oxidized, producing sulphuric acid, and the bottle will become clear. This is, on a small scale, the process already described for the manu-

facture of the acid. To acidify the water to any great extent, the operation of igniting the sulphur, &c., must be repeated several times.

Ex. h.—Sometimes the sulphur and nitre are mingled together before mixing. Mix some small pieces of sulphur with one-eighth their own weight of saltpetre or nitre, and place them on a piece of lead or other metal in a shallow vessel filled with water. Ignite the sulphur mixture, and invert a gas jar over it while burning. The acid fumes that arise will be condensed by the water, forming an impure sulphuric acid. This liquid, as well as that resulting from the previous experiments, may be tested by blue litmus paper, the reddening of which will shew the presence of an acid; or by adding to some of the water in a wine-glass a few drops of chloride of barium, when a white precipitate of sulphate of barium will be formed, quite insoluble in water or in nitric or hydrochloric acids. The formation of this precipitate is a sure test of the presence of sulphuric acid, either free or in a state of combination.

22. The reaction occurring between the sulphurous acid, the nitric acid, the steam and the air, during the process of manufacturing sulphuric acid, may be explained as follows :—

1. Three atoms of sulphurous acid, $3SO_2$, abstract in the first instance two atoms of oxygen, $2O$, from each atom of nitric acid, N_2O_5, forming with the steam three atoms of sulphuric acid, $3(H_2SO_4)$, and leaving two atoms of binoxide of nitrogen, or nitric oxide, $2NO$.
2. This nitric oxide (as shewn *Chap. VIII.* § 16) instantly absorbs oxygen from the air, and is converted into hyponitric anhydride, $2NO_2$, giving rise to its characteristic orange-coloured vapours.
3. Each atom of hyponitric anhydride so formed, parts with half its oxygen to two atoms of sulphurous acid, $2SO_2$, forming with the steam as before two atoms of sulphuric acid, $2H_2SO_4$. The resulting nitric oxide, $2NO$, again takes oxygen from the air, and again gives it to the sulphurous acid, and this alternate action goes on so long as the supply of sulphurous acid, air and steam continues.

In this way a small amount of nitric acid suffices to convert a large quantity of sulphur*ous* into sulphuric acid; the necessary oxygen being supplied by the air, the nitrous acid acting as a carrier between it and the sulphurous acid.

The several stages of the process may be thus expressed by symbols, the second and third forming a constantly recurring series:—

1st. $N_2O_5+3SO_2+3H_2O = 2NO+3H_2SO_4$.

2nd. $2NO+2O = 2NO_4$.

3rd. $2NO_2+2SO_2+2H_2O = 2NO+2H_2SO_4$.

23. Sulphuric acid in its most concentrated state is a colorless, inodorous, oil-like liquid, exceedingly corrosive, having an intensely acid taste, and a great affinity for water. It freezes at —15°, and boils at 620°.

So great is the affinity of sulphuric acid for water, that a bottle three-parts filled with strong acid, left open for twenty-four hours in a damp atmosphere, would be found completely filled.

24. When sulphuric acid is suddenly mixed with water, much heat is evolved, and on cooling, condensation is found to have taken place, the two liquids occupying less space than before mixture. Whenever two bodies are thus condensed, or have their particles driven closer together, heat is produced.

Ex. i.—Mix four parts by measure of strong sulphuric acid with one part of water in a thin glass flask, it will in a few moments become too hot to hold.

Ex. k.—If an ounce of sulphuric acid, of the temperature of 32o, be poured over an *ounce* of snow or pounded ice, of a like temperature, the density of the combined substances will be greater than that of the two substances separately; and in this condensation so much latent heat will be given out, that the mixture will reach a temperature, almost equal to that of boiling water.

Ex. l.—Pulverize quickly in a mortar one *pound* of ice, and pour over it in a tumbler one ounce of sulphuric acid, previously cooled to 32°. Stir the mixture, and the whole will become fluid. If a ther-

mometer be immersed in it, the temperature will be found very nearly 0°, or 32° below the freezing point.

These two experiments may be reconciled by observing, that in the latter case the quantity of acid is only sufficient to liquefy the ice, which, on leaving the solid state, abstracts heat from surrounding objects, while in the former case the acid being in greater proportion, not only liquefies the ice, but unites with the water as in *Ex. i.*

Ex. m.—The corrosive action of sulphuric acid on organic bodies is accompanied by a blackening or charring effect, owing to its abstracting from them the elements of water, and leaving their carbon behind. A drop of strong sulphuric acid let fall on linen or paper speedily destroys it. A stick dipped into the strong acid is soon blackened. Its effects on starch, gum, and sugar have been already noticed. (*Chap. IV. Ex. n. and q.*)

25. Sulphuric acid is extensively used in the manufacture of soda from common salt; also in making chlorine, in dyeing, calico-printing, gold and silver refining, and in the purification of oil and tallow. In short, its chemical uses are almost innumerable.

26. Sulphur forms two compounds with hydrogen, the former of which is a very useful and important body: they are—

Sulphuretted Hydrogen . . . $H_2S = 34$.
Persulphuret of Hydrogen, . . $HS = 33$.

27. **Sulphuretted Hydrogen,** H_2S, is a colorless, transparent gas, having an extremely disagreeable taste and smell; it is somewhat heavier than air, and soluble in about one-third its bulk of water, to which it imparts its own offensive odor and taste.

It does not support combustion, but is itself combustible, burning with a blue flame and producing water and sulphurous acid when there is an abundant supply of air, but depositing sulphur when the oxygen is deficient.

This gas is highly poisonous to animals. "A small bird dies immediately in air containing $\frac{1}{1500}$th part of sulphuretted hydrogen; $\frac{1}{1000}$th part will kill a middle-sized dog; and $\frac{1}{250}$ part will kill a horse."—*Brande.*

28. Sulphuretted hydrogen is produced by the putrefaction of all organic substances containing sulphur, as flesh, eggs, &c.; it arises from vegetable matter decomposing in connexion with water and gypsum (sulphate of lime); it frequently issues from the earth in volcanic districts, and occurs in small quantities in certain mineral springs, as at Harrogate and elsewhere.

29. Sulphuretted hydrogen is usually procured by the action of hydrochloric acid on sulphuret of antimony, or by the decomposition of sulphuret of iron by sulphuric acid.

The first method yields the purest gas, and may be carried out by placing finely powdered sulphuret of antimony, Sb_2S_3, in a retort or flask, pouring over it an equal weight of strong hydrochloric acid, HCl, and heating the mixture; the gas should be collected over *warm* water, *and if possible in the open air.* The reaction is as follows:—

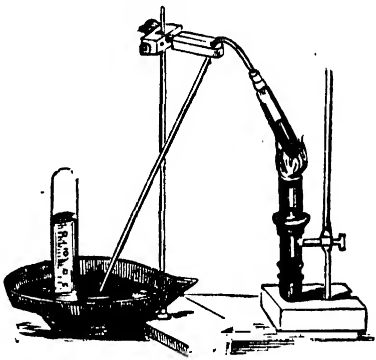


Fig. 50.

$Sb_2S_3 = 340$	$\begin{cases} 3S = 96 \\ 2Sb = 244 \end{cases}$	$3H_2S$	$= 102$
$6HCl = 219$	$\begin{cases} 6H = 6 \\ 6Cl = 213 \end{cases}$	$2SbCl_3$	$= 457$
559	559		559

$$Or,\ Sb_2S_3 + 6HCl = 2SbCl_3 + 3H_2S.$$

The second method, which yields the gas most readily, is by putting protosulphuret of iron, FeS, into a generating bottle, and pouring over it an equal weight of dilute sulphuric acid. Each atom of iron (diatomic) in the protosulphuret (FeS) replaces the double-atom of hydrogen in the acid (H_2SO_4), and forms protosulphate of iron ($FeSO_4$); the double-atom of hydrogen unites with the sulphur, and escapes as sulphuretted hydrogen (H_2S). The reaction will be seen by the following diagram:—

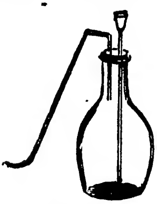


Fig. 51.

$FeS = 88 \begin{cases} S = 32 \\ Fe = 56 \end{cases}$		$H_2S = 34$
$H_2SO_4 = 98 \begin{cases} H_2 = 2 \\ SO_4 = 96 \end{cases}$		$FeSO_4 = 152.$
186 186		186

Or thus, $FeS + H_2SO_4 = FeSO_4 + H_2S$.

30. The solution of sulphuretted hydrogen in water is of great use to the chemist, as a test for metals, and more especially, it is an excellent substance for discovering minute portions of lead, with which it gives a black colored precipitate; with antimony it gives an orange, with arsenious acid a bright yellow, and with zinc, a white precipitate. These precipitates are sulphurets of the respective metals.

31. The solution may be easily made by passing a current of the gas through water for a short time; in the open air if possible.

Ex. n.—Arrange a series of two or more flasks as in the figure, place some water in the smaller one B, and some sulphuret of iron in the larger one A. Pour through the funnel a little diluted sulphuric acid, sulphate of iron will be formed, and the sulphuretted hydrogen will pass off into the second flask, and be absorbed by the water and so on through a series of flasks, if necessary.

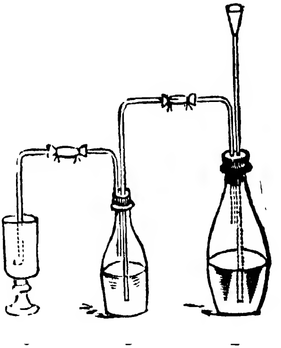


Fig. 52.

32. The solution of sulphuretted hydrogen in water is so prone to decomposition, by the oxygen of the air, that it soon spoils. The simplest and readiest mode of applying the gas is by having a small generating flask, like fig. 51, always at hand, supplied with water and sulphuret of iron; on the addition of a few drops of sulphuric acid, the gas will be immediately evolved when wanted, and passed at once through the delivery tube into the vessel containing the solution to be tested, as in C, fig. 52.

Ex. o.—Pass some gas as just described into a solution of acetate of lead or nitrate of silver in water, black sulphuret of lead or silver will be formed; this accounts for the blackening of white-lead paint when exposed to coal-gas, which is almost always found to contain sulphuretted hydrogen; silver goods are tarnished and blackened from a similar cause. The discoloring of a silver spoon when used in eating an egg, arises from the action of the sulphur which the egg contains.

Ex. p.—Treat a solution of tartar emetic (tartrate of antimony and potass) in the same way, orange-colored sulphuret of antimony will be precipitated.

Ex. q.—Perform the like experiment with sulphate of zinc, white sulphuret of zinc will fall. Hence, white paint made from oxide of zinc, keeps its color much better than that from oxide of lead.

These experiments may be tried with different metallic solutions, and the color of the sulphurets observed.

33. **Selenium,** Se.—This is an exceedingly rare sub-

stance, sometimes found associated with sulphur, which element it greatly resembles in its chemical properties. It is a reddish brown solid body, semitransparent, and having an imperfect metallic lustre.

When heated in the air it exhales an odor like decaying horse-radish.

34. Selenium combines with oxygen and hydrogen, forming—

Selenious acid, H_2SeO_3, corresponding to Sulphurous acid, H_2SO_3.

Selenic acid, H_2SeO_4, corresponding to Sulphuric acid, H_2SO_4.

Seleniuretted hydrogen, H_2Se, corresponding to Sulphuretted hydrogen, H_2S.

"Selenic acid has one property possessed by no other simple acid—namely, that of dissolving gold."—*Wilson.*

Seleniuretted hydrogen is a colorless gas, excessively irritating, to the nostrils, air passages, and lungs, even when respired in very minute quantity, and producing all the symptoms of a violent cold.

EXERCISES ON CHAPTER XIV.

1. Give the symbol, combining equivalent, and physical characteristics of sulphur.
2. Where does sulphur occur pure?
3. Where is sulphur found in a state of combination?
4. State the effects of various temperatures on the appearance of sulphur.
5. What is the allotropic form of sulphur?
6. Distinguish between *sublimation* and *distillation.*
7. How may the purity of sulphur be ascertained?
8. What are the chief compounds of sulphur and oxygen?
9. What is the product of the combustion of sulphur in the air?
10. State the properties of sulphurous acid.
11. How may sulphurous acid be liquefied?
12. Explain the process for obtaining sulphurous acid gas.
13. Give the formula for the above process.

14. How may aqueous sulphurous acid be obtained?
15. What are the different forms of sulphuric acid?
16. How may the name *oil of vitriol* be accounted for?
17. Whence does Nordhausen acid derive its name, and for what is it used?
18. Explain the process for obtaining Nordhausen acid.
19. How is anhydrous sulphuric acid obtained, and what are its properties
20. What is hydrated sulphuric acid, and how may it be obtained from sulphurous acid gas?
21. Describe the process for manufacturing the sulphuric acid of commerce. What is the best test for sulphuric acid?
22. Describe the chemical changes that take place during the process, noting each step symbolically.
23. What are the leading properties of sulphuric acid?
24. What effect is produced by mixing sulphuric acid and water?
25. What are the chief uses of sulphuric acid?
26. Give the names, symbols, and equivalents for the compounds of sulphur and hydrogen.
27. What are the properties of sulphuretted hydrogen?
28. Name some natural sources of sulphuretted hydrogen
29. How may sulphuretted hydrogen be artificially prepared? Give formulæ for each method.
30. In what way is a solution of sulphuretted hydrogen a test for different metals?
31. How may an aqueous solution of sulphuretted hydrogen be made?
32. How can sulphuretted hydrogen be most readily applied as a test?
33. State the nature and properties of selenium.
34. What compounds does selenium form with oxygen and hydrogen? State some of their properties.

CHAPTER XV.

PHOSPHORUS AND ITS COMPOUNDS.

Symbol, P. *Combining weight*, 31.

1. **Phosphorus** is a solid body, of a waxy appearance, easily cut, transparent and colorless when fresh made, but turning yellow by exposure to light. It melts at 108°, and at 550° is converted into vapor. It is insoluble in water,

but may be dissolved by ether, alcohol, bisulphuret of car bon, and oils.

2. When exposed to the air it undergoes a slow combustion, combining with oxygen and emitting a whitish smoke, (phosphorous anhydride PO_3) which is luminous in the dark. Hence its name *phosphorous* or *light-bearer*, from *φῶς*, *light*, and *φερειν*, *to bear*.

3. Phosphorus is highly inflammable, being easily set on fire by friction, hence its use in making lucifer* matches.

It sometimes takes fire spontaneously in warm weather, and should therefore be kept under water, and be at all times handled with great caution, since the heat of the hand is sufficient to inflame it when dry, and a burn from it is generally very severe.

4. Phosphorus never occurs in nature in a free or uncombined state, but is very widely diffused in union with oxygen, under the form of phosphoric acid, H_3PO_4, and thus forms an ingredient both of vegetable and animal bodies. The phosphoric acid in these cases is *generally* in combination with lime or magnesia.

5. The bones † and shells of animals contain phosphorus; in fact, their strength and stiffness is owing to the phosphates of lime and magnesia. Phosphorus is found also in the blood, the animal excretions, the brain, liver, nerves, &c.; also in eggs, oysters, onions, wheat, milk, flour, &c.

6. Phosphorus is chiefly obtained from bones. Bones consist of gelatine, lime, and phosphoric acid. The phosphorus is procured from the decomposition of the phosphoric acid, as follows—

The bones are first calcined or burned to whiteness, to remove the

---

* *Lucifer*, from *lux*, *lucis*, light, and *ferre*, to bring, is the exact Latin equivalent for the Greek, *phosphorus*.

† The skeleton of a man weighs from 10 to 12 pounds, and contains from 1½ to 2 pounds of phosphorus.

gelatine. The ashes consisting chiefly of phosphate of lime, $3Ca, 2PO_4$, are then ground and mixed with about two-thirds their weight of dilute sulphuric acid. The sulphuric acid combines with the lime forming an insoluble sulphate, which is removed by filtering. The *filtrate*, that is, the liquid which passes through the filter, is an impure solution of phosphoric acid, H_3PO_4: this is evaporated to the consistence of treacle, then mixed with charcoal powder, and strongly heated in an iron vessel to remove all moisture.

The dried mass is then placed in a retort made of stoneware or fire-clay, A, and exposed to a white heat. The carbon takes the oxygen of the phosphoric acid, passing off at D, as carbonic oxide. The phosphorus distils over in vapor, and passes through the tube B, and may be condensed below the surface of water, in the vessel C. The tube B should be kept warm lest the phosphorus should condense in it, and block up the passage.

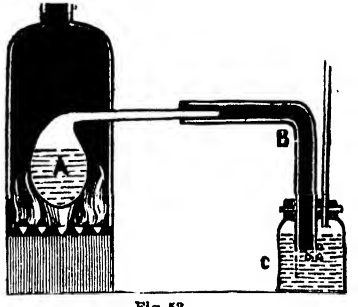


Fig. 53.

The decomposition of the phosphoric acid may be thus shewn in symbols:

$$2(H_3PO_4) + 5C = 3H_2O + 5CO + 2P.$$

The phosphorus is melted a second time under hot water in cylindrical moulds, so that it is generally seen in the form of long pencil-like sticks.

7. A singular modification of this element has been discovered by Professor Schrötter, called *allotropic* or *amorphous* phosphorus. It is a reddish brown substance, which is luminous and inflammable only at high temperatures. It shines in the dark when heated to 393°, and at 500° returns into the condition of ordinary phosphorus. From the fact of the amorphous variety being less combustible than common phosphorus, and not giving rise (at low temperatures) to the same poisonous vapors, it has been proposed to use it in the manufacture of lucifer matches.

8. Phosphorus combines readily with oxygen, forming the following compounds, the last of which is, perhaps, the most important :—

	Anhydride,		Acid.
Hypophosphorous .	P_2O	;	H_3PO_2.
Phosphorous* .	P_2O_3	;	H_3PO_3.
Phosphoric . .	P_2O_5	;	H_3PO_4.

9. **Phosphorous Anhydride,** P_2O_3.—This is formed by slow combustion of phosphorus in a *dry* atmosphere; or by burning that element in a limited supply of air. The substance so formed is anhydrous, and appears as a white powder, which, when exposed to air or moisture, changes first, into phosphorous, and then, into phosphoric acid.

10. **Phosphorous acid,** $3H_2O,P_2O_3$, or H_3PO_3, may be readily prepared by adding water to the teriodide or terchloride of phosphorus, PI_3 or PCl_3. Mutual decomposition takes place, the hydrogen of the water unites with the iodine or chlorine to form hydriodic or hydrochloric acids, while the oxygen combines with the phosphorus; by heating the liquid the hydrogen acid, HI or HCl, together with most of the water, is driven off, while the phosphorous acid combined with another portion of water, remains.

$$2PI_3 + 6H_2O = 6HI + 3H_2O,P_2O_3 = 6HI + 2(H_3PO_3).$$
$$2PCl_3 + 6HO = 6HCl + 3H_2O,P_2O_3 = 6HCl + 2(H_3PO_3).$$

Ex. a.—Place a small piece of phosphorus (dried between some blotting paper) into a test tube, and cover it with some iodine, the two bodies will combine, forming the teriodide of phosphorus, PI_3, as in *Ex. d.* page 16. When the action is over, add some water to dissolve the mass, and then heat it to evaporate the hydriodic acid; the hydrated phosphorous acid which remains will, on cooling, become a solid crystalline mass.

11. **Phosphoric Acid,** H_3PO_4. This substance has been already noticed, § 6, as being produced by the action of

* The existence of this oxide has been denied by some chemists; the red substance that is so termed being considered by them to be the amorphous phosphorus previously described.

sulphuric acid on the phosphate of lime contained in bone ashes. It is the constant product of the combustion of phosphorus in oxygen, or in a plentiful supply of common air, appearing as a white smoke, which condenses into snow-like flakes of anhydrous phosphoric acid, P_2O_5.

Ex. b.—For the combustion of phosphorus in air and in oxygen, see *Ex. a. Chap. VIII.*, and *Ex. c. Chap. IV.*

Ex. c.—Perform similar experiments in *dry* jars or bottles, the anhydrous phosphoric acid will condense on the sides.

Ex. d.—Collect some of the dry powder of the last experiment, put it quickly into a dry watch-glass, and add a few drops of water, or drop some of it at once into water; the acid and water will rapidly combine, with a hissing noise, great heat, and sometimes even a violent explosion.

12. Phosphoric acid has an intensely sour taste, and reddens vegetable blues; it is not poisonous, at least, in small quantities. Although not generally corrosive, it must, when taken medicinally, be sucked through a quill or glass tube to keep it from corroding the teeth, which it would do very rapidly, from its power of dissolving the phosphate of lime which they contain.

E. e.—Place a thin slice of phosphorus on a piece of blotting-paper, and sprinkle over it some flowers of sulphur or some powdered charcoal, after a while the phosphorus will spontaneously inflame. This is owing to the rapid absorption of oxygen by the porous powder, and its imparting it again to the phosphorus, while from its being a bad conductor of heat, the cooling of the latter is prevented.

Ex. f.—Put a fragment or two of phosphorus into a glass, *b*, nearly fill the glass with hot water, and then direct a stream of oxygen or common air from a bag or bladder on to the melted phosphorus, as in the diagram. The phosphorus burning in presence of the oxygen will present the appearance of *fire under water*.

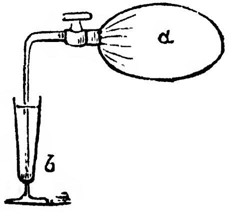


Fig. 54.

Ex. g.—Put a piece of phosphorus, about the size of a pea, into a phial with a quarter of an ounce of ether; cork the phial, and let it remain a few days, occasionally shaking it. About a grain of the phosphorus will be dissolved, and the solution, which is sometimes

called *phosphorised ether*, will, if exposed to the air, be luminous in the dark.

Ex. h.—Phosphorised oil may be made by carefully heating a grain of phosphorus in a test tube with a teaspoonful of olive oil; the solution should be kept in a well-corked phial, and in the dark. A few drops of phosphorized oil or ether may be rubbed on the hands or face without danger (unless there be any wounds on them); the parts so moistened will appear luminous in the dark.

Ex. i.—Moisten some blotting-paper with a few drops of phosphorized ether, or solution of phosphorus in bisulphuret of carbon, and hang it up to dry. The moisture will speedily evaporate and leave the phosphorus on the paper in a state of very minute division, in which state it attracts oxygen so energetically, that it ignites spontaneously.

Ex. k.—Pour a few drops of phosphorized ether on a lump of loaf sugar, and place the sugar in hot water; the ether and phosphorus will rise in vapor and take fire on the surface as they come into contact with the oxygen of the air. The effect is increased if the surface of the water is made to undulate by blowing on it.

13. **Phosphuretted Hydrogen, $H_3P = 34$.** This is a colorless, transparent gas, of a very disagreeable odor, and highly poisonous. The nauseous smell of decaying animal matter is partly owing to the presence of this gas. The singular phenomenon of the will-o'-the wisp, or jack-o'-lantern, where a flame is seen by night moving over marshy ground, is believed to be due to the presence of this gas, which is spontaneously inflammable.

14. Phosphuretted hydrogen may be obtained by heating some pieces of phosphorus with caustic potass or lime, and water, in a test tube as in Fig. 1, or in a small flask with a long tube as in the annexed diagram. The flask, &c. should be completely filled with water, and the delivery end of the tube be also under water. On the application of heat the following reaction takes place: four atoms of phosphorus decompose six atoms of water, producing three atoms of hypophosphorous acid, H_3PO_2, and one of phosphu-

Fig. 55.

retted hydrogen, PH_3, which, when thus prepared, is spontaneously inflammable, each bubble as it escapes into the air taking fire, and changing into a ring of smoke, which widens as it rises, and consists of phosphoric acid and watery vapour.

The hypophosphorous acid unites with the potass or lime. to form an hypophosphite of those substances. Neither the lime nor the potass is decomposed, but seems merely by *its presence* to bring about the decomposition, which may be thus expressed symbolically:

$$4P + 6H_2O + 3KHO = 3KPH_2O_2 + 3H_2O + H_3P.$$

Ex. l.—Put some pieces of zinc and phosphorus in a flask or test glass, and pour over them some dilute sulphuric acid. The water will be decomposed as in *Chap. V.* § 14, but the hydrogen, as it is set free, will combine with the phosphorus and form inflammable phosphuretted hydrogen; it will not, however, be so pure as by the former method.

Ex. m.—Throw some pieces of phosphuret of calcium into a glass of water, phosphuretted hydrogen will be given off and inflame as before.

Ex. n.—If the hydrated phosphorous acid obtained in § 10 be heated in a small retort it will be decomposed, yielding hydrated phosphoric acid and phosphuretted hydrogen, which, however, when thus procured, ceases to be spontaneously inflammable.

$$4\,(H_3PO_3) = 3(H_3PO_4) + H_3P.$$

EXERCISES ON CHAPTER XV.

1. State the properties of phosphorus.
2. Whence does phosphorus derive its name?
3. Why is it used for lucifer matches?
4. Where is phosphorus found in nature?
5. Name some substances which contain phosphorus.
6. How is phosphorus obtained?
7. What is amorphous phosphorus.
8. What are the compounds of phosphorus with oxygen?
9. What is phosphorous acid?
10. How may phosphorous acid be prepared?
11. How is phosphoric acid obtained?
12. What are the properties of phosphoric acid?
13. What is phosphuretted hydrogen?
14. How may phosphuretted hydrogen be obtained?

CHAPTER XVI.

CHLORINE, BROMINE, IODINE, AND FLUORINE.

1. These four bodies, Chlorine, Bromine, Iodine, and Fluorine, form a small natural group of elementary bodies, closely resembling each other in their properties, and widely differing from the elements already noticed, inasmuch as they have a much greater affinity for hydrogen than for oxygen, and form with it the hydrogen-acids already named, *Chap. VII.*, § 8.

2. The symbols and combining proportions of these bodies are as follow :—

Chlorine	Cl	=	35.5.
Bromine	Br	=	80.
Iodine	I	=	127.
Fluorine	F	=	19.

3. **Chlorine** (Cl) is a gaseous body, discovered by a German chemist named Scheele in 1774, but first shown to be a distinct element by Sir Humphrey Davy, who gave it its present name *chlorine,** on account of its yellowish green color.

4. Chlorine exists abundantly in nature, both in plants and animals, but especially in the mineral kingdom ; it is generally found in combination with the metal *sodium,* forming *chloride of sodium* or common salt, NaCl.

5. Chlorine possesses a powerful and disagreeable odor, and is excceedingly suffocating, even when largely diluted with air. It is readily absorbed by water, and this liquid, especially if recently boiled, will take up about twice its bulk of gas, and acquire the same color, taste and smell.

* Chlorine, from χλωρος, *chlōros*, yellowish green.

6. Chlorine is only gaseous at common temperatures and pressures, for if exposed to great cold or to a pressure of four atmospheres (60lbs. to the square inch), it condenses into a limpid, yellow liquid, about one-third heavier than water. In its gaseous state it is about $2\frac{1}{2}$ times as heavy as air.

7. Chlorine is most readily prepared by heating together in a flask or retort equal weights of hydrochloric acid and black oxide of manganese.

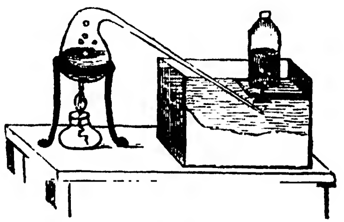


Fig. 56.

The gas as given off may be collected at the pneumatic trough, either over *hot* water, or over a strong solution of salt and water. Being heavier than common air, it may be collected by *displacement*, by carrying the delivery tube to the *bottom* of the bottle in which the gas is to be received, as in fig. 48, omitting the water from the bottle there drawn. The chlorine gradually rises and expels or displaces the air. The bottles should in either case be removed as soon as filled, and the stoppers well greased a little before they are inserted.

8. In this decomposition four atoms of hydrochloric acid, HCl, are required for one of binoxide of manganese, MnO_2; two atoms of water and one of protochloride of manganese are formed, while one molecule of chlorine is set free; thus—

$$4HCl + MnO_2 = 2H_2O + MnCl_2 + Cl_2$$
$$146 + 87 = 36 + 126 + 71.$$

9. Another mode of obtaining chlorine, especially when required in large quantities, is by the action of sulphuric acid on a mixture of black oxide of manganese and common salt. When sulphuric acid is added to a mixture of binoxide of manganese and chloride of sodium, NaCl, or

common salt, sulphate of sodium and sulphate of manganese are formed and chlorine gas given off.

Two atoms of sulphuric acid $2(H_2SO_4)$ are decomposed.

(i) The two atoms of hydrogen in one atom of acid are replaced by two atoms of sodium, forming sulphate of sodium (Na_2SO_4).

(ii.) The two atoms of hydrogen in the other atom of acid are replaced by one atom of manganese (a di atomic element) forming sulphate of manganese ($MnSO_4$).

(iii.) The four atoms of hydrogen displaced unite with the two atoms of oxygen, to form two atoms of water; while

(iv.) The two atoms of chlorine originally combined with the sodium are disengaged in the form of gas; thus:—

$$2NaCl + MnO_2 + 2(H_2SO_4) = Na_2SO_4 + MnSO_4 + 2H_2O + 2Cl.$$

10. Chlorine has a great affinity for hydrogen and the metals; with the former it produces hydrochloric acid; with the latter, compounds termed chlorides. Its combinations with oxygen are exceedingly unstable, that is, they are easily decomposed, while with nitrogen it forms one of the most explosive bodies known.

Hydrochloric acid	HCl	=	36.5
Hypochlorous acid	$HClO$	=	52.5
Chlorous acid .	$HClO_2$	=	68.5
Chloric acid .	$HClO_3$	=	84.5
Perchloric acid .	$HClO_4$	=	100.5
Terchloride of azote	NCl_3	=	120.5

11. Chlorine is a powerful antiseptic and disinfectant, it is also largely employed as a bleaching agent, either in the gaseous state, or in solution in water, or in combination with lime (bleaching powder, commonly called chloride of lime.)* It entirely destroys animal and vegetable coloring

* *Chloride* of lime is a misnomer, for the *element* chlorine cannot unite with the *compound* lime. The substance is properly hypochlorite of calcium.

matter, but only when moisture is present, for *dry* chlorine gas has no bleaching power.

12. The disinfecting and bleaching properties of chlorine are owing to its strong affinity for hydrogen. As a destroyer of contagion it acts by decomposing sulphuretted hydrogen and other noxious vapors in the atmosphere. As a bleaching instrument it acts *indirectly* by decomposing water, and combining with the hydrogen to form hydrochloric acid, while the nascent* oxygen unites with, and removes the coloring matter, and is the real bleaching agent.

Ex. a.—Pour a solution either of litmus, indigo, carmine, turmeric, or ink, into a small bottle of chlorine; do not put in the stopper, but cover the mouth with a piece of flat glass, and then shake it well, the coloring matter will speedily disappear.

Ex. b.—Suspend a nosegay of various-colored flowers in a jar as in Fig. 47; put into the saucer or plate a little bleaching powder, and add some hydrochloric acid diluted with warm water; chlorine will be evolved, and the flowers will soon lose their color and become perfectly white.

Ex. c.—The great affinity of chlorine for hydrogen, and its corresponding indifference to carbon may be seen, by immersing a well-lighted candle in a jar of the gas, when it will burn with a red smoky flame, the chlorine combining with the hydrogen of the combustible, while the carbon appears as smoke.

Ex. d.—This may be still more strikingly shewn by dipping a rag or a piece of folded paper into turpentine and immersing it, while moist, into a jar of the gas. The turpentine, C_5H_4, will ignite spontaneously, and the jar will be filled with black smoke and soot. Here,

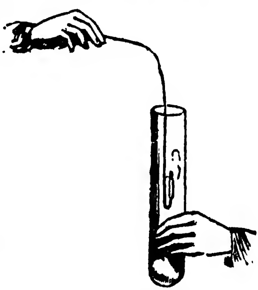


Fig. 57

* Nascent, see chap. iii. § 25


Fig. 58.

again, the chlorine unites with the hydrogen, and rejects the carbon.*

Ex. e.—Repeat *Ex. h.* chap. xii., the result will be the same as in the last two experiments.

Ex. f.—Place a piece of phosphorus in a deflagrating spoon, fig. 7, and immerse it in a jar of chlorine, it will take fire and burn with a feeble light, forming a great contrast to that produced by its combustion in oxygen. This feebleness of light is owing to the product of combustion (chloride of phosphorus) being a liquid, and not a solid as in the case of phosphoric acid. See chap. v. § 16.

Ex. g.—Drop into a jar of chlorine some filings of antimony, heated to about 80°, rapid scintillating combustion will take place, attended by a *white* flame; the product of combustion will be chloride of antimony, $SbCl_3$. "Similar experiments may be performed with most of the other metals, giving various colored flames. Silver and lead give a white color; zinc and tin, a bluish white; iron, a vivid red; copper, a dull red; arsenic, gold, and tellurium, a green. These metals must be heated previously to immersion in chlorine, and although a very low degree of heat is sufficient for some, as lead and zinc, yet the combustion is much more vivid if they be made still hotter; for instance, the more fusible metals so hot that they can be just held in the fingers without burning them, and the more refractory metals till they change color by heat. The jars used should be rather tall and narrow, for those experiments where the powdered metal is sprinkled in."—*Francis.*

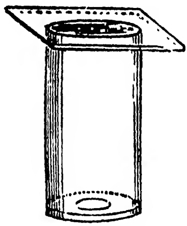


Fig. 59.

14. **Hydrochloric acid,** HCl, is a colorless, transparent *gas,* having intensely acid properties; it is exceedingly soluble in water, which at ordinary temperatures absorbs more than four hundred times its bulk of the gas. It is this aqueous solution which is sold in the shops, and is sometimes called *muriatic acid.*

* The chlorine may be obtained for this and similar experiments by putting a little "chloride of lime" into a wide-mouthed bottle or jar, and adding some dilute hydrochloric acid, the gas will very soon fill the jar.

14. Hydrochloric acid is generally obtained by decomposing common or sea-salt by sulphuric acid, hence it was formerly named *marine acid* or *spirit of salt*.

$$2NaCl + H_2SO_4 = Na_2SO_4 + 2HCl.$$

Hydrochloric acid may also be formed by the direct union of equal measures of hydrogen and chlorine.

Ex. h.—Mix together in a jar equal measures of chlorine gas and hydrogen; expose the mixture to light, the gases will unite, and form the hydrochloric acid in a gaseous state, and without water.

If the above be exposed to the direct rays of a hot sun, the gases will ordinarily combine with such rapidity as to explode with violence. "Procure a tube 12 inches long, of half an inch internal diameter, fill it with the mixed gases, and expose it to a full light, the combination of the gases will almost instantly be seen to commence by the cloudy appearance produced within the tube; now cover over the tube, and the action will cease until a second time exposed, and thus by repeating the experiment, the action of light upon the gases is beautifully shown, while from the small size of the tube, there is no danger in submitting the whole to the direct rays of the sun."—*Francis*.

Ex. i.—Perform *Ex. g*, chap. xii. with equal measures of chlorine and hydrogen, the gases will explode with a loud report. Turn the mouth of the bottle downwards into a basin of water colored blue by litmus. The hydrochloric acid gas will be rapidly absorbed, the water will rise in the bottle, and become red, proving the acid character of the gas.

15. When hydrochloric acid is added to a metallic oxide, mutual decomposition ensues, the oxygen and chlorine change places, for the hydrogen of the acid combines with the oxygen of the oxide to form water, while the chlorine combines with the metal to form a metallic chloride.

All the hydracids act on metallic oxides in a similar way, the hydrogen being displaced by the metal.

16. **Bromine**, Br.—This is one of the rarer elements, being found only in small quantities, combined with sodium or magnesium, in sea-water, in marine plants and animals, and in some mineral springs, especially those of Kreuznach in Prussia. It was first discovered in 1826, by

M. Balard, a French chemist, who obtained it from the liquid remaining after the crystallization of common salt at the salt-works of Montpelier, and named it *Bromine** on account of its very offensive and suffocating odor.

17. Bromine is a highly poisonous and very corrosive liquid, of a reddish brown color, and about three times as heavy as water. It is very volatile, emitting even at ordinary temperatures a red vapor, resembling that of nitrous acid, and nearly $5\frac{1}{2}$ times as heavy as atmospheric air. When cooled a little below 0°, bromine becomes solid, crystalline, and brittle. At 116° it boils. It is slightly soluble in water, but more so in alcohol and ether; like chlorine, it is a powerful bleaching agent. When added to a solution of starch, it forms a compound of a pale orange color.

18. Bromine unites with oxygen and hydrogen to form bromic and hydrobromic acids, its combinations with the other elements are termed bromides. The bromide of potassium is occasionally used in medicine and the arts.

Bromic acid	$HBrO_3$	=	129
Hydrobromic acid	HBr	=	81

Ex. i.—Place a piece of phosphorus in a deflagrating spoon and immerse it in a jar of the vapor of bromine; it will immediately burst into flame.

Ex. k.—Immerse a burning taper into a jar of the vapor of bromine; the flame will assume a dull red and green tint, and be speedily extinguished.

Ex. l.—Drop into a jar of the vapor of bromine a few grains of powdered tin, or some filings of antimony, the particles of metal will become red-hot, and resemble a shower of fire.

"The best method of performing these experiments is to have a tall jar. Let three or four drops of bromine fall to the bottom of the jar, and pour a spoonful of boiling water upon the bromine; or better, touch it with the point of a red-hot wire."—*Francis.*

19. **Iodine**, I.—This element was discovered by M. Courtois in 1812. It is found, like bromine, though in greater

* Bromine, from βρῶμος, *bromos*, a strong smell.

quantity, in sea-water, marine plants and animals, and in some mineral springs.

20. Iodine is generally obtained from the ashes of sea-weed (*kelp*), by dissolving them in water, filtering and evaporating the solution so as to separate the salt by crystallization, then mixing the remaining liquid with sulphuric acid and binoxide of manganese, and gently heating the mixture in a retort; the iodine rises in violet-colored vapors, which may be condensed in a receiver in a solid form. The reaction is similar to that shown in the preparation of chlorine from common salt; the iodide of sodium in the seaweed, &c., taking the place of the chloride: thus—

$$2NaI + MnO_2 + 2(H_2SO_4) = Na_2SO_4 + MnSO_4 + 2H_2O + \mathbf{2I}.$$

21. Iodine at ordinary temperatures, is a solid, occurring usually in the form of crystalline plates or scales, of a bluish-black color, and metallic lustre, somewhat resembling plumbago, and nearly 5 times as heavy as water. It is slightly volatile at low temperatures, at 225° it fuses, and at 375° it boils, giving off its characteristic violet-colored vapor, which is the heaviest aeriform substance known, being more than eight times as heavy as common air. (sp. gr. 8.716.)

22. Iodine has a peculiar, pungent odor, and an acrid taste; it stains the skin of a deep yellowish-brown color, which, however, soon disappears. It is readily dissolved by alcohol, but so slightly by pure water, that a pound of water will not dissolve more than a grain of it. One of its most characteristic properties is the production of a fine blue compound with starch (iodide of starch).

Ex. m.—Place a little iodine at the bottom of a flask or long test tube, and carefully heat it over a lamp, it will be converted into a violet-colored vapor, while the upper part of the tube will be covered with black, lustrous crystals of the condensed or sublimed iodine.

Note the density of the vapor by the sluggishness with which it moves as the tube is inclined to one side or the other.

Ex. n.—Place a few small pieces of phosphorus, on a tile or plate of metal, and when dry, sprinkle a little iodine over them. The two

elements will combine and give out sufficient heat to ignite the phosphorus.

Ex. p.—Put 20 grains of iodine into an ounce of water, scarcely any will be dissolved. Add 30 grains of iodide of potassium, and the whole will be dissolved immediately.

This aqueous solution, as well as the solution in alcohol, is used in medicine, being found useful in cases of glandular swellings, especially *scrofula* and *goître*.

The following notice of the discovery and application of iodine is from Hoblyn's "Manual of Chemistry." "The practical application of chemical knowledge in the discovery of *new* substances is full of interest. One instance may here suffice. In the manufacture of soap, the vessel employed is found to be corroded; a scientific chemist analyses the corroding matter, and the result is the discovery of one of the most singular and important chemical elements, *iodine*. Curiosity is excited; the origin of the new substance is traced to the sea plants, from whose ashes soda, the principal ingredient of soap, is obtained, and ultimately to the sea-water itself. It is thence hunted through nature, discovered in salt-mines and springs, and pursued into all bodies which have a marine origin; among the rest, into sponge. A medical practitioner then calls to mind a reputed remedy for the cure of one of the most grievous and unsightly disorders to which the human species is subject—the *goître*, which infests the inhabitants of mountainous districts to an extent that, in this favored land we have happily no experience of, and which was said to have been originally cured by the ashes of burnt sponge. Led by this indication, he tries the effect of iodine on that complaint, and the result establishes the extraordinary fact, that this singular substance, taken as a medicine, acts with the utmost promptitude and energy on *goître*, dissipating the largest and most inveterate in a short time, and acting as a specific, or natural antagonist, against that odious deformity. The history of chemistry is full of facts of equal, or greater interest and importance."

Ex. o.—Make a little starch paste, such as laundresses use, by boiling some starch with water, in a test tube. Put a small quantity of this paste into a glass of water, and add a few drops of a solution of iodine, a deep blue iodide of starch will be immediately produced The blue color disappears when the liquid is heated.

23. The chief compounds of iodine with oxygen, hydrogen, and nitrogen are—

Iodic acid	HIO_3	=	176
Hydriodic acid	HI	=	128
Teriodide of Nitrogen	NI_3	=	395

24. **Teriodide of Nitrogen,** NI_3, is an explosive substance like the terchloride, NCl_3, but more easily prepared, and by no means so dangerous.

Ex. p.—It is obtained by putting iodine into a solution of ammonia, NH^3. The ammonia is decomposed, hydriodic acid is formed by the union of its hydrogen with one portion of the iodine, while its nitrogen unites with another portion; the result is an insoluble dark brown powder, which may be collected and allowed to dry of itself in small portions, on separate pieces of blotting paper. When dry, this powder explodes at the slightest touch, so that it cannot be moved. If only *small* quantities are operated with, and *ordinary care* be used, there is no danger attending its explosion.

25. **Fluorine,** F.—This is an elementary substance contained in a few minerals, and especially in Derbyshire spar or fluor-spar, which is a fluoride of calcium, CaF. It is also found in small quantities in sea water, and in various plants; also in shells, coral, and bones, and especially in the enamel of the teeth.

26. Fluorine has never yet been obtained perfectly pure, owing to its great affinity for most of the metals and for silicon, (the basis of flint, and one of the ingredients of glass). It is believed to be a yellowish-brown gas, similar in its general properties to chlorine.

27. **Hydrofluoric acid,** *symbol* HF, *comb. wt.* 19. This is the most important of the compounds of fluorine. It is obtained from fluor-spar by the action of sulphuric acid, just as hydrochloric acid is obtained from salt.

$$CaF_2 + H_2SO_4 = CaSO_4 + 2HF.$$

It is a colorless, corrosive gas, distinguished for its power of dissolving glass, hence it must be prepared in platinum or leaden vessels. Advantage is sometimes taken of this corrosive property, for the purpose of engraving or etching on glass.

Ex q.—"A plate of glass is covered with a thin coating of wax and blackened by holding it over the flame of a candle, the design being then traced upon the surface with a needle. The plate thus prepared, is placed over a leaden vessel of sufficient size, contain-

ing a mixture of pulverized fluor-spar and sulphuric acid, which is gently warmed. The pungent, acid-smelling vapor of hydrofluoric acid (HF) is evolved, and attacks the glass wherever it is bare. After 10 or 20 minutes, the plate is removed, and gently warmed, in order to free it from wax, when the etching becomes distinctly visible. The vapors of hydrofluoric acid are, it must be remembered, very pernicious, and attack even the skin; the greatest care is therefore required."—*Schoedler*.

Fig. 60.

EXERCISES ON CHAPTER XVI.

1. Why may Chlorine, Bromine, Iodine and Fluorine be classed together?
2. Give their symbols and combining numbers.
3. By whom was chlorine discovered? Whence does it derive its name?
4. Where is chlorine found in nature?
5. State some of the leading properties of chlorine.
6. How may chlorine be liquefied?
7. How may chlorine be prepared from hydrochloric acid?
8. Explain fully the reaction that takes place between hydrochloric acid and binoxide of manganese.
9. How may chlorine be obtained from salt? Describe the reactions that take place.
10. What are the principal compounds of chlorine? Give their symbols and combining numbers.
11. For what is chlorine chiefly used?
12. How may the disinfecting and bleaching properties of chlorine be accounted for?
13. State the properties of hydrochloric acid.
14. Describe the methods of obtaining hydrochloric acid.
15. What change takes place when hydrochloric acid is added to a metallic oxide?
16. Where is bromine found? by whom was it discovered? whence does it derive its name?
17. State the leading properties of bromine.
18. What are the chief compounds of bromine?
19. When and by whom was iodine discovered?
20. Describe the mode of preparing iodine.
21. State the leading properties of iodine.
22. What is the most characteristic test of iodine?
23. Name some of the chief compounds of iodine, with their symbols and combining proportions.

24. Describe the properties and mode of preparation of teriodide of azote.
25. Where is fluorine met with in nature?
26. What is known about fluorine?
27. State the properties and mode of preparation of hydrofluoric acid.

CHAPTER XVII.

BORON AND SILICON.

Boron B = 11
Silicon Si = 28

1. **Boron** is one of the rarer elements, being but sparingly diffused through nature, and then only in combination with oxygen, forming boracic anhydride, B_2O_3, from which substance it may be obtained, by fusion with potassium, thus—

$$B_2O_3+6K = 3K_2O+2B.$$

2. Boron is a brownish-green, infusible, and insoluble powder, unalterable by exposure to the air, at common temperatures; when heated to about 600°, it combines with oxygen and burns with considerable brilliancy, forming boracic anhydride, B_2O_3

3. **Boracic acid**, HBO_2. This substance is found in a free state in the hot springs of Tuscany, and in combination with soda, forming *borax* or biborate of sodium ($Na_2B_4O_7+10H_2O$), which is brought to England from Eastern Asia under the name of *tincal*.

4. Boracic acid is generally seen in the form of white pearly crystals, soluble in water and alcohol. The latter solution burns with a beautiful green flame.

When heated, boracic acid first parts with its water of crystallization, and then fuses, forming when cold, a transparent, colourless glass-like substance, B_2O_3.

Ex. a.—Break a crystal or two of borax to powder, and add a few drops of sulphuric acid, to set free the boracic acid; next pour on the mass a little alcohol, and set light to it, the green color already referred to will be readily observed.

$$(Na_2B_4O_7 + 10H_2O) + H_2SO_4 = Na_2SO_4 + 4(H_3BO_3) + 5H_2O.$$

5. **Silicon,** Si, never occurs in nature in a free or uncombined state, but is always found in union with oxygen, forming *silicic acid* or *silica,* SiO_2; it may be obtained from the silico-fluoride of potassium by fusion with excess of potassium; thus—

$$K_2SiF_6 + 4K = 6KF + Si.$$

6. Silicon* seems to occupy among mineral substances a similar place to that held by carbon among vegetable and animal bodies, being a chief constituent of the crust of the globe, and, perhaps, next to oxygen, the most abundant element in nature. It is a dark brown powder, insoluble in water, and not easily oxidized; when strongly heated, it combines with oxygen, forming silicic acid.

7. **Silicic anhydride,** or silica, SiO_2, may be prepared by heating flint or rock crystal to redness, and quenching it in water, when it may be easily reduced to a fine, white, tasteless, gritty powder which is nearly pure silica, and which after being strongly heated is quite insoluble in water, or acids, except the hydrofluoric.

8. Quartz and rock-crystal are nearly pure silica, whilst flint, agate, cornelian, chalcedony, onyx, jasper, opal, amethyst, and many other minerals consist of silicic acid, more or less combined with metallic oxides. These minerals are distinguished by a peculiar hardness, which enables them to give sparks with steel, and to scratch glass.

9. Silica neither tastes sour, nor reddens vegetable blues, yet when raised to a high temperature it exhibits powerful acid properties, and combines with many metallic oxides, forming the class of compounds termed silicates. The

* Silicon or silicium, from the Latin, *silex, silicis,* flint or sand, of which it is the basis.

minerals just named are silicates; glass, porcelain, and earthenware are silicates of potass, soda, lime, or alumina formed at a high heat.

10. If silica (SiO_2) be heated to redness with an *excess* of potass, soda, or lime, it forms a soluble silicate, which may be decomposed by adding a stronger acid, such as the hydrochloric, the **Silicic acid** is separated in an amorphous state under the form of a gelatinous or jelly-like mass, soluble in pure water. It loses its solubility, however, when heated. Its formula is H_4SiO_4.

11. Silica in its soluble state is nearly always present in springs and mineral waters: in this state, too, it enters into plants, in many of which it exists abundantly, particularly in the stalks of the various grasses, to which it communicates stiffness and strength; it may be readily detected in their ashes after they are burned.

12. The soluble glass formed in § 10 has been used to wash over timber, and other combustible materials to render them fire-proof; the effect is to fix a coating of silica on the timber, &c., which in some degree preserves it, yet too slightly to be of much service.

EXERCISES ON CHAPTER XVII.

1. Where is boron found, and how may it be obtained?
2. Give the symbol, combining equivalent and leading properties of boron.
3. What is boracic acid, and where is it found?
4. Describe the properties of boracic acid.
5. How may silicon be obtained?
6. Whence does silicon derive its name? What are its chief properties?
7. What is silica, and how may it be obtained?
8. Name some natural and artificial compounds of silica.
9. Why is silica considered an acid?
10. What is the allotropic form of silica, and how may it be prepared?
11. Of what use is silica in nature?
12. For what purpose has "soluble glass" been employed?

CHAPTER XVIII.

THE EQUIVALENT POWER OR ATOMICITY OF THE ELEMENTS; ACID-TYPES.

1. By the term atomicity or equivalent power of the elements is meant their power of combining with, or replacing different proportions of hydrogen.

2. The atomicity of elements which do not combine with hydrogen is ascertained by their power of combining with, or replacing, chlorine.

3. The term *equivalent* was formerly used as synonymous with atomic or combining weight, but it is now employed to denote *that proportion of a substance which will replace in a compound body a given proportion of another substance.*

4. The equivalent number may or may not be the same as the combining proportion, or atomic weight; for one, two, three, or even four atoms of one element may be replaced by a single atom of another.

5. The first column of the table on page 9 gives the equivalent power of the elements, that is, their power of combining with or replacing *one*, *two*, *three*, or *four* atoms of hydrogen.

6. Elements which can only replace one atom of hydrogen, are termed **Mono-atomic**; those which can replace two atoms are called **Di-atomic**; three atoms, **Tri-atomic**; and four atoms, **Tetr-atomic**.

7. Thus, an atom of hydrochloric acid (HCl) consists of one atom of hydrogen (H) combined with one atom of chlorine (Cl), and is, in fact, an atom of chloride of hydrogen.

Now the atom of H in the atom of HCl may be replaced by an atom, of either sodium (Na), potassium (K), silver (Ag), iodine (I), &c., giving rise to chlorides of those several elements, NaCl, KCl, AgCl, ICl, &c.

Or the atom of Cl in the atom of HCl, may be replaced by an atom of either iodine, bromine, fluorine, &c., giving rise to the hydracids, HI, HBr, HF, &c.

Similarly the atom of Cl in the atom of NaCl, KCl, AgCl, above named, can be replaced by an atom of either I, Br, F, &c., forming NaI, NaBr, NaF, &c., as the case may be.

These several elements, which by single atoms are equivalent to each other in their saturating power, are termed **Mono-atomic.**

8. In comparing an atom of water (H_2O) with an atom of Hydrochloric acid (HCl) we see that the atom of oxygen has the power of saturating *twice* as much hydrogen as the atom of chlorine does.

So, also, we find that the hydrogen (H_2) in *two* atoms of hydrochloric acid (2HCl) can be replaced by one atom of calcium (Ca), copper (Cu), Lead (Pb). &c., forming $CaCl_2$, $CuCl_2$, $PbCl_2$, &c. in place of H_2Cl_2.

Hence these and similar elements which have double the equivalent power of hydrogen are called **Di-atomic.**

9. If an atom of gaseous ammonia (NH_3) be compared with the atom of hydrochloric acid (HCl), it will be seen that the atom of nitrogen has the power of saturating *three* times as much hydrogen as the atom of chlorine does.

So, also the hydrogen (H_3) in *three* atoms of hydrochloric acid (3HCl) can be replaced by one atom of bismuth (Bi), or gold (Au), &c., forming $BiCl_3$, $AuCl_3$, &c., in place of H_3Cl_3.

The Nitrogen, Bismuth and Gold, having three times the equivalent power of hydrogen, are termed **Tri-atomic.**

10. On comparing marsh-gas (CH_4) with hydrochloric acid (HCl), we find the one atom of carbon (C), saturating *four times* as much hydrogen as the atom of chlorine does.

Or, again, the hydrogen (H_4) in *four* atoms of hydrochloric acid

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So, also, we find that the hydrogen (H_2) in *two* atoms of hydrochloric acid (2HCl) can be replaced by one atom of calcium (Ca), copper (Cu), Lead (Pb). &c., forming $CaCl_2$, $CuCl_2$, $PbCl_2$, &c. in place of H_2Cl_2.

Hence these and similar elements which have double the equivalent power of hydrogen are called **Di-atomic.**

9. If an atom of gaseous ammonia (NH_3) be compared with the atom of hydrochloric acid (HCl), it will be seen that the atom of nitrogen has the power of saturating *three* times as much hydrogen as the atom of chlorine does.

So, also the hydrogen (H_3) in *three* atoms of hydrochloric acid (3HCl) can be replaced by one atom of bismuth (Bi), or gold (Au), &c., forming $BiCl_3$, $AuCl_3$, &c., in place of H_3Cl_3.

The Nitrogen, Bismuth and Gold, having three times the equivalent power of hydrogen, are termed **Tri-atomic.**

10. On comparing marsh-gas (CH_4) with hydrochloric acid (HCl), we find the one atom of carbon (C), saturating *four times* as much hydrogen as the atom of chlorine does.

Or, again, the hydrogen (H_4) in *four* atoms of hydrochloric acid

($4HCl$), can be replaced by one atom of tin (Sn), or one atom of platinum (Pt), &c., forming $SnCl_4$, $PtCl_4$, &c., in place of H_4Cl_4.

Hence Carbon, Tin, Platinum, having four times the saturating or equivalent power of hydrogen, are termed **Tetr-atomic.**

11. One atom of a mono-atomic element is equivalent to and can replace one atom of another mono-atomic element.

One atom of a di-atomic element can replace one atom of another diatomic element, or two atoms of a mono-atomic element; and so on.

The equivalent or replacing powers of the elements may be shewn thus:—

Mono-atomic, $E^{i} = E^{i}$.

Di-atomic, $E^{ii} = E^{ii}$ *or* $E^{i} + E^{i}$.

Tri-atomic, $E^{iii} = E^{iii}$ *or* $E^{ii} + E^{i}$ *or* $E^{i} + E^{i} + E^{i}$.

Tetr-atomic, $E^{iv} = E^{iv}$ *or* $E^{iii} + E^{i}$ *or* $E^{ii} + E^{ii}$ *or* $E^{ii} + E^{i} + E^{i}$ *or* $E^{i} + E^{i} + E^{i} + E^{i}$.

12. Not only do the elements combine with each other, but there are several groups of elements which act collectively as a single element; such groups have been termed compound radicals.

13. These compound radicals vary in atomicity or equivalent power, being mono-atomic, di-atomic, &c., like the simple elements.

14. It has been already stated (Chap. VII., § 12) that all the acids may be regarded as compounds of hydrogen; they may, however, be conveniently classed according to the **type**, or model, on which they seem constituted.

15. The simplest class of acids is formed (as in the case of Hydrochloric acid), on the type of a molecule of hydro-

gen (HH), by replacing one of the atoms of hydrogen by another mono-atomic element (Cl), thus:—

Hydrogen, 2H; Hydrochloric Acid, HCl.

$$\left.\begin{matrix} H \\ H \end{matrix}\right\} \qquad \left.\begin{matrix} H \\ Cl \end{matrix}\right\}$$

16. The second class of acids is formed, as in the case of Nitric acid, on the type of *one* atom of water (H_2O), by replacing *one* of the atoms of hydrogen by a mono-atomic radical (NO_2), thus:—

Water, H_2O; Nitric Acid, HNO_3.

$$\left.\begin{matrix} H \\ H \end{matrix}\right\} O, \qquad \left.\begin{matrix} H \\ NO_2 \end{matrix}\right\} O.$$

17. A third class of acids (*dibasic*) is formed (as in the case of sulphuric acid), on the type of two atoms of water ($2H_2O$), by replacing *two* atoms of hydrogen by one of a di-atomic radical (SO_2), thus:—

2 Water, $2H_2O$; Sulphuric Acid, H_2SO_4.

$$\left.\begin{matrix} H_2 \\ H_2 \end{matrix}\right\} O_2. \qquad \left.\begin{matrix} H_2 \\ SO_2 \end{matrix}\right\} O_2.$$

18. A fourth class of acids (*tribasic*) is formed (as in the case of phosphoric acid) on the type of *three* atoms of water ($3H_2O$), by replacing *three* atoms of hydrogen by one of a tri-atomic radical (PO), thus:—

3 Water, $3H_2O$; Phosphoric Acid, H_3PO_4.

$$\left.\begin{matrix} H_3 \\ H_3 \end{matrix}\right\} O_3. \qquad \left.\begin{matrix} H_3 \\ PO \end{matrix}\right\} O_3.$$

19. In a similar way the oxides and hydrated oxides of the different metals may be regarded as formed on the type of one or more atoms of water, by the replacement of the hydrogen by one or more atoms of the metal according to its atomicity. Thus:—

CHAPTER XIX.

METALS.

1. The second and larger group of elementary bodies consists of the metals. They are fifty-two in number, seven only of which—viz.: gold, silver, mercury, copper, iron, tin, and lead—were known to the ancients. All the others, except three,—antimony, zinc, and bismuth*—have been discovered within the last 150 years. Some of the metals are widely diffused through nature, while many are of very rare occurrence.

2. The metals are distinguished from the non-metallic elements, by a peculiar brilliancy, termed the *metallic lustre*, and by their being good conductors of heat and electricity. They exhibit, however, great variations in these as well as in their other properties, and in their colors. When reduced to a state of minute division, as by filing, or any other means, the metallic lustre is lost, but the color remains.

3. With the single exception of mercury, the metals are solid at the usual temperature of the air; but they all *fuse* or become liquid by heat, and if the temperature could be raised sufficiently high, they would probably all pass off into vapor, as zinc, mercury, and arsenic are known to do.

4. The fusing or melting point of metals varies extremely from —39° F., at which temperature mercury appears solid, to the intense heat of the oxy-hydrogen blow-pipe, which is necessary to fuse platinum and some other metals.

* Antimony was described by Basil Valentine in the fifteenth century. Zinc is noticed by Agricola in 1520 A.D. Bismuth is first mentioned by Paracelsus in the sixteenth century.

The following table shewing the fusibility of different metals is taken from Thenard's Chemistry, Vol. 1.

		F.	
	Mercury	—39°	
Fusible below a red-heat 1000°	Potassium	136	
	Sodium	190	
	Tin	442	
	Cadmium, about	442	
	Bismuth	497	
	Tellurium	500	
	Lead	612	
	Zinc	773	
	Antimony, a little below red heat		
Infusible below a red heat	Silver	1873	
	Copper	1996	
	Gold	2016	
	Cobalt, Nickel — rather less fusible than iron		
	Iron (cast)	2786	
	Iron (wrought)	3300	
	Manganese	3300	
	Palladium, Molybdenum, Uranium, Tungsten, Chromium	almost infusible, and not to be procured in buttons by the heat of a smith's forge.	Fusible by the oxy-hydrogen blow-pipe.
	Titanium, Cerium, Osmium, Iridium, Rhodium, Platinum, Tantalum	Infusible in the heat of a smith's forge.	

5. Among the volatile metals, that is, those which pass off in vapor at comparatively moderate temperatures, are mercury, cadmium, arsenic, tellurium, zinc, potassium and sodium.

6. The metals differ greatly from each other in specific gravity. Some, like potassium and sodium, are lighter than

water; the more common ones are seven or eight times as heavy as water. The heaviest metal is platinum.

Table of the Specific Gravities of the more common metals, at 60° F. compared with Water as the standard.

Potassium	·865	Iron (cast)	7·2
Sodium	·972	Iron (wrought)	7·7
Magnesium	1·7	Nickel	8·2
Aluminium	2·6	Copper	8·8
Arsenic	5·8	Bismuth	9·8
Antimony	6·7	Silver	10·5
Zinc (cast)	6·8	Lead	11·4
Manganese	7·0	Mercury	13·5
Zinc (rolled)	7·1	Gold	19·3
Tin	7·2	Platinum	21·5

7. Some of the metals are malleable,* that is, they can be rolled into thin sheets or beaten into leaves, as gold, copper, tin, &c. Other metals are so brittle that they may be reduced to powder by hammering, as arsenic, antimony, &c.

8. Nearly all the malleable metals are ductile,† that is, they can be drawn out into wire. It will be seen from the following table, that the ductility and malleability of the same metal, are not always in proportion to each other.

Order of Malleability	*Order of Ductility.*
Gold.	Gold.
Silver.	Silver.
Copper.	Platinum
Tin.	Iron.
Platinum.	Copper.
Lead.	Zinc.
Zinc.	Tin.
Iron.	Lead.

* Malleable, from the Latin *malleus*, a hammer, maul or mallet.

† Ductile, from the Latin *ducere*, to draw or lead.

Gold is the most malleable of the metals. A grain of gold may be beaten out so as to cover above fifty square inches of surface, or somewhat more space than is occupied by the two pages of this book now open. Gold is also the most ductile of the metals. A gold wire may be drawn out so fine, that 550 feet of it will weigh only one grain.

10. The tenacity* of metals is measured by finding the greatest weight which can be supported by a wire of a certain thickness. The following table shews the weights capable of being supported by wires of different metals, each wire being about one-tenth of an inch (·840 of a line) in diameter. From this table it would appear that iron is the most tenacious of the metals.

	lbs.
Iron . . .	549·25
Copper . .	302·278
Platinum . .	274·32
Silver . .	187·237
Gold . .	158·753
Zinc . .	109·54
Tin . . .	34·63
Lead . .	27·621

11. All bodies expand when heated, but metals expand the most. An iron wire heated to 212° will be about $\frac{1}{800}$th part longer than it was at 32°. The action of the common thermometer depends on the ready expansibility of mercury.

Ex. a.—The expansion of metals may be shewn by arranging a brick, a knitting-needle, or piece of wire, and a thin board, as in the figure. On heating the wire with a spirit lamp, the board if carefully balanced will be overturned.

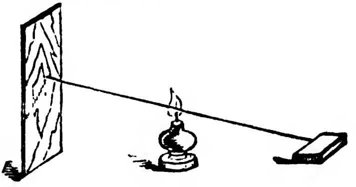


Fig. 61.

12. Many of the metals combine with

* Tenacity, from the Latin *tenax*, *tenacis*, holding on.

each other, forming compounds, termed *alloys;* if one of the metals be mercury, the compound metal is called an *amalgam*. Thus brass is an alloy of copper and zinc, while a looking-glass is "silvered" at the back with an amalgam of mercury and tin.

Ex. b.—Melt one ounce of zinc in a ladle, and add to it an ounce of lead, and an ounce of bismuth. They will form an alloy, so fusible that it may be melted in hot water.

Ex. c.—An alloy of eight parts bismuth, five lead, and three tin, will melt at the heat of boiling water, or 212°.

Ex. d.—If two parts of mercury be added to the last when melted, an amalgam will be formed, which remains fluid at a much lower temperature, and is used for injecting anatomical preparations.

13. All the metals combine more or less readily with oxygen, and frequently in more proportions than one. When a metal unites with oxygen it loses its metallic lustre, and takes a dull, earthy appearance. The metallic oxides may be conveniently arranged in three classes;—1st, Acidifiable oxides;—2nd, Basic oxides or bases;—3rd, Neutral oxides.

14. **Acidifiable oxides** are those which, in combination with water produce acids (Chap. vii. § 10), that is, compounds having for the most part a sour taste, and the property of changing vegetable blues to red, and browns to yellow; their chief and only essential characteristic, however, being their power of combining with *bases* to form *salts*.

15. **Bases**, or basic oxides are chiefly distinguished by their power of combining with and neutralizing acids. They include the *alkalies*, which are *strong** bases, having alkaline properties, that is, a peculiar acrid taste, such as that of soda, or lime, and the power of turning vegetable blues to green, and yellows to a reddish brown.

* Acids or bases are spoken of as *strong* or *weak* according as their acid or basic properties are more or less strongly marked.

16. The following is a list of the principal **basic** oxides :—

ALKALIES.

Potass or Oxide of Potassium		K_2O
Soda ,, ,, ,, **Sodium**		Na_2O
(Ammonia ,, ,, **Ammonium)**		$(NH_4)_2O$

ALKALINE EARTHS.

Baryta or Oxide of Barium	BaO
Strontia ,, ,, **Strontium**	SrO
Lime ,, ,, ,, **Calcium**	CaO
Magnesia ,, ,, **Magnesium**	MgO

NON-ALKALINE BASES.

Protoxide	**of Iron**	FeO
Sesquioxide ,,	,,	Fe_2O_3
Oxide ,,	**Zinc**	ZnO
Suboxide ,,	**Copper**	Cu_2O
Protoxide ,,	,,	CuO
,, ,,	**Lead**	PbO
,, ,,	**Silver**	Ag_2O
Suboxide ,,	**Mercury**	Hg_2O
Protoxide ,,	,,	HgO
,, ,,	**Tin**	SnO
Sesquioxide ,,	**Chromium**	Cr_2O_3
Protoxide ,,	**Nickel**	NiO
,, ,,	**Cobalt**	CoO
,, ,,	**Manganese**	MnO
Sesquioxide ,,	,,	Mn_2O_3
,, ,,	**Antimony**	Sb_2O_3

17. **Neutral Oxides** are combinations of oxygen which do not possess the properties either of acids or bases.

Water, or oxide of hydrogen, **H_2O—Carbonic Oxide,** or oxide of carbon, CO,—and **Nitrous oxide,** or laughing gas, **N_2O,—are** examples of neutral oxides among the **metalloids.**

Binoxide of Manganese, MnO_2,—and **Binoxide of Lead**, PbO_2,—and **Sesquioxide of Bismuth**, Bi_2O_3,—are examples of neutral oxides among the metals.

18. The *non-metallic* oxides are mostly *acidifiable;* hence, sulphuric, carbonic, nitric and phosphoric acid.

The *metallic* oxides are for the most part *bases;* as, potass, soda, lime, oxide of lead, &c.

The strong bases are all oxides, or compounds containing *one* atom of oxygen, and its equivalent of metal.

The weaker bases are usually sesqui-oxides.

The metallic acids are found only among the highest oxides, while the indifferent or neutral oxides hold an intermediate position.

19. The oxides of manganese afford a striking illustration of the gradual change of properties by the increase in the proportion of oxygen: thus

Protoxide of Manganese	MnO	strong base.
Sesquioxide ,, ,,	Mn_2O_3	weak base.
Binoxide ,, ,,	MnO_2	neutral.
Manganic acid	H_2O+MnO_3	strongly acid.
Permanganic acid	$H_2O+Mn_2O_7$	

20. The chief metallic acids are—

Arsenious acid	H_3AsO_3	**Antimonious acid**	$H_2Sb_2O_5$
Arsenic ,,	H_3AsO_4	**Antimonic** ,,	$H_2Sb_2O_6$
Chromic ,,	H_2CrO_4	**Manganic** ,,	H_2MnO_4
Molybdic ,,	H_2MoO_4	**Permanganic** ,,	$HMnO_4$
Titanic ,,	H_2TiO_4	**Tungstic** ,,	H_2WO_4

To these may be added the following, which, under certain circumstances, exhibit acid properties:—

Ferric acid	or	**Teroxide of Iron**	FeO_3
Auric ,,	,,	**Sesquioxide of Gold**	Au_2O_3
Stannic ,,	,,	**Peroxide of Tin**	

21. **The metals differ greatly in their relative affinities for oxygen.**

Potassium and sodium, for instance, are oxidized by mere exposure to the air, and decompose water at all temperatures whenever they come into contact with it.

Iron and copper remain unchanged in dry air, and do not decompose water at common temperatures; but each of them oxidizes slowly in a moist atmosphere, and combines rapidly with oxygen when heated to redness.

Iron, however, seems to have a stronger affinity for oxygen than copper; for the former can decompose water at a red heat, while the latter cannot produce that effect.

Mercury will remain unoxidized even in a moist atmosphere, and, though it combines with oxygen at about 650° F., its oxide is reduced to a metallic state when heated to redness.

Neither gold, nor platinum, can be oxidized by exposure to heat alone. They are only made to combine with oxygen by an indirect and difficult process.

22. The metals may be conveniently arranged in six groups, based on the nature of the compounds which they form with oxygen.

Many of the metals being exceedingly rare, and not sufficiently important to be considered in an elementary work, their names have been omitted from the following table, but will be found in the general list of elements, pages 9, 10, 11.

Classification of the Metals.

1. *Metals of the Alkalies.*

Potassium.	Lithium.
Sodium.	(*Ammonium*).

2. *Metals of the Alkaline Earths.*

Barium.	Calcium.
Strontium.	Magnesium.

3. *Metals of the Earths proper.*

Aluminium.	Glucinium, &c.

4. *Metals whose oxides form powerful bases.*

Manganese.	Zinc.
Iron.	Cadmium.
Chromium.	Copper.
Nickel.	Lead.
Cobalt.	Bismuth.

5. *Metals whose oxides are weak bases or acids.*

Tin.	Antimony.
Arsenic.	Tungsten, &c.

6. *Metals whose oxides are reducible by heat alone.*

Mercury.	Gold.
Silver.	Platinum, &c.

EXERCISES ON CHAPTER XIX.

1. How many metallic elements are there? Name those which were known to the ancients.
2. How are the metals to be distinguished from the non-metallic elements?
3. What is the natural state of the metals at ordinary temperatures?
4. Give the fusing point of the principal metals.
5. Name some of the volatile metals.
6. State the specific gravities of some of the metals.
7. When are metals said to be malleable?
8. Why are some metals termed ductile?
9. Name the metals in their order (1.) of malleability, (2.) of ductility.
10. How is the tenacity of metals estimated? Which metal is the most tenacious?
11. How may the expansibility of metals by heat be illustrated?
12. What are alloys and amalgams?
13. How may the metallic oxides be classified?
14. What are acids?
15. What are bases? What are the properties of an alkaline base?
16. Name the most important bases, and state their composition.
17. What are neutral oxides? Give examples.

18. How does the proportion of oxygen affect the character of metallic oxides?
19. Illustrate this by the series of oxides of manganese.
20. Name the chief metallic acids, and state their composition.
21. Illustrate the different affinities of metals for oxygen.
22. How may the metals be conveniently classified?

CHAPTER XX.

METALS OF THE ALKALIES.

POTASSIUM, SODIUM, LITHIUM.

Potassium.	*Symbol*	K;	*combining weight*	39
Sodium.	,,	Na;	,, ,,	23
Lithium.	,,	L;	,, ,,	7
(*Ammonium.*)	,,	(NH_4)	,, ,,	(18)

1. **Potassium,** (or Kalium, whence the symbol K.) is a silvery white metal, so soft at ordinary temperatures that it may be readily cut with a knife. At 32° it is brittle, and at 150° it fuses. It is considerably lighter than water, its specific gravity at 60° being 0.865. It occurs abundantly in nature, but always in a state of combination.

2. The most striking chemical property of potassium is its great affinity for oxygen, with which it unites to form the alkali potass or potash. The affinity of potassium for oxygen is so strong, that if the metal be exposed to the air it immediately becomes covered with a film of oxide, hence it can only be preserved by being kept under naphtha, a liquid which contains no oxygen.

3. Potassium was first obtained by Sir H. Davy in 1807, by decomposing *hydrate of potassium,* KHO, by means of a voltaic current. The oxygen both of the water and the potass appeared at the positive pole, while the hydrogen of

the water appeared with the metal potassium at the negative pole.

4. Potassium is now generally procured, either by fusing hydrate of potassium, KHO, with iron turnings, Fe; or by heating together carbonate of potass, K_2CO_3, and charcoal, C.

In the first case the hydrate of potassium is decomposed by the heated iron, which combines with the oxygen, while the liberated potassium sublimes, and may be collected in a cool part of the apparatus.

$$2KHO+2Fe=2FeO+2H+2K.$$

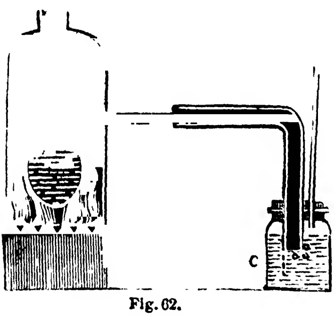


Fig. 62.

In the second process, which is similar to that for obtaining phosphorus, (page 105), the charcoal combines with the oxygen, forming carbonic oxide, and the potassium is set free as before.

The vessel C must be filled with naphtha, and kept quite cool.

$$K_2CO_3+2C=3CO+2K.$$

Ex. a.—Throw a globule of potassium upon a little water in a plate, the water will be immediately decomposed. Its oxygen combines with part of the potassium to form potass, while its hydrogen unites with another portion of the metal forming potassiuretted hydrogen, which burns with a rose-colored flame.

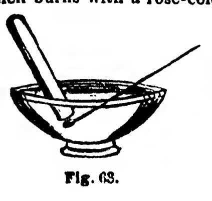


Fig. 63.

Ex. b.—Put a globule of potassium on the surface of ice, a similar action will take place, and a similar light will be produced.

Ex. c.—Perform the experiment described in Chapter IV., p. 21.

Ex. d.—Perform *Ex. a.* with water previously colored yellow by turmeric or rhubarb, or rendered blue by infusion of violets or red-cabbage; the yellow will be changed to brown and the blue to green, shewing that an alkali has been formed.

Ex. e.—Put a small fragment of potassium with a little flour of sulphur into a test tube, and hold it for an instant or two in the flame of a spirit lamp. The sulphur and potassium will instantly combine, forming sulphuret of potassium. Light and heat will be evolved, and the tube will be broken.

Ex. f.—Repeat the preceding, using iodine instead of sulphur; iodide of potassium will be formed, attended with a brilliant light, and the breakage of the tube, as before.

5. Every substance in which oxygen is known to be present may be decomposed by the metal potassium, and almost all such decompositions are accompanied by the disengagement of light and heat. Hence, though not applied to any important uses in the arts, it is a valuable agent in the hands of the chemist for analysing other bodies.

6. **Potass,** Potassa, or Protoxide of Potassium, $K_2O, = 94$.—Pure or anhydrous Potass may be obtained by heating potassium in *dry* air or oxygen, whereby the metal becomes converted into a white solid, which possesses powerful *alkaline* properties, and is extremely caustic; it fuses at a red heat, and is volatile at a white one.

When this solid is moistened with water, it slakes like quick-lime, evolving great heat, and forming the hydrate of potassium, described in the next paragraph.

7. **Hydrate of Potassium,** $KHO, = 56$.—This substance, generally known as caustic potass, is formed whenever a solution of potass in water is evaporated to dryness. Thus when potassium is burned on the surface of water (as in *Ex. a.*), the water will be found, after the combustion, to contain potass; if this solution be evaporated to dryness, the solid substance that remains is not pure potass, but a hydrate of potassium, each molecule of which consists of one atom of water and one of potass ($K_2O + H_2O = 2KHO$). No increase

of temperature can remove the water from this compound, for it may be fused and even volatilized unchanged.

8. Hydrate of potassium may be procured by boiling together hydrate of calcium, CaO,H_2O, and carbonate of potassium, K_2CO_3; the carbonic acid leaves the potass, and uniting with the lime, forms an insoluble carbonate of calcium, $CaCO_3$, which is precipitated; the clear liquid is then drawn off, and evaporated to dryness in an iron or silver vessel, when the hydrate of potassium is left as a white, soluble, deliquescent solid.

$$K_2CO_3 + CaO,H_2O = CaCO_3 + 2KHO.$$

If the heat be continued, the mass fuses, and may then be cast into moulds, forming pencil-like sticks, in which form it is much used by surgeons as a caustic, and by chemists for removing carbonic acid gas from gaseous mixtures.

9. Potass is a constant ingredient in all fertile soils, and is found in almost all plants, especially in the leaves and young wood, hence it has been termed the *vegetable* alkali. It is known in commerce under the name of *pot-ash* and *pearl-ash*, which are mixtures of the hydrate and carbonates of potass, more or less impure.

The terms *pot-ash* and *pearl-ash* point to the source whence we obtain potass, namely, the *ashes* of land plants.

Ex. g.—Pour a little water over some ashes of burned wood in a basin, stir the mixture, and leave it a little while in order to dissolve what is soluble. When the heavier parts have settled, strain off the clear liquor, boil it to a thick syrup, and set it by to cool; when cold, it will be found a solid mass of impure carbonate of potass. If this be boiled with lime as already described, hydrate of potassium (caustic potass) may be obtained.

If the hydrate be wanted very pure, it may be dissolved in alcohol (spirits of wine), which dissolves only the pure hydrate, leaving earthy and other impurities behind; the alcholic solution may then be strained off, and the alcohol removed by distillation in a platinum or silver vessel. Metallic vessels must be used, as potass combines with silica (flint) and alumina (clay), and would consequently corrode glass or porcelain retorts.

10. Potass possesses powerful alkaline properties, being

intensely acrid, and neutralizing the strongest acids. It is extremely *caustic*, or destructive of animal matter, whether living or dead, hence its use in surgery.

It is much employed in the manufacture of glass, also in making soft-soaps, and in the processes of dyeing and calico-printing.

Ex. h.—Pour a few drops of oil on some water in a glass, and then stir in a little potass, the potass combines with the oil, forming a soapy mixture, which, if the water were evaporated, would become what is commonly called *soft-soap*.

This accounts for the soft, greasy feel of potass when touched with the fingers, the skin is decomposed and the potass forms a soap with its oily elements. For a like reason soiled clothes are cleansed by boiling with potass, for the alkali unites with the acid of the grease contained in them, and forms a soap, soluble in water.

Ex. i.—Mix in a crucible 10 parts of potass (common pearl-ash), 15 of sand, and 4 of charcoal, and heat it in a strong fire; the potass and silica will combine and form a glass (silicate of potassium) which is soluble in water.

11. **Sodium**, (or Natrium, whence the symbol Na), the metallic base of the alkali Soda, has a silvery lustre, and a slightly reddish tinge. It is so soft that it may be cut with a knife, but, unlike potassium, it retains this softness at the freezing point, 32°; it melts at 194°, and goes off in vapor at a red heat. It is somewhat lighter than water, its specific gravity at 60° being 0.972.

12. Sodium very much resembles potassium in its properties, and, like that metal, occurs abundantly in nature, but always in a state of combination. Its most important compounds are carbonate of sodium, Na_2CO_3, and chloride of sodium, NaCl.; in the latter form it composes vast beds of rock-salt in various parts of the world, but is still more extensively diffused in the waters of the ocean. It is also found very plentifully in sea-weeds and marine plants, and seems to be always present in the bodies of animals.

12.* Sodium has a great affinity for oxygen, and must, therefore, like potassium, be kept in naphtha.

It decomposes water rapidly, but does not burst into flame, unless it be placed on a bad conductor of heat,—or its movements be retarded by mixing starch or gum in the water,—or unless the water be heated; in the latter case it takes fire at once, burning with a characteristic *yellow* flame, and forming a solution of soda.

Ex. k.—Try experiments with sodium similar to those described with potassium, *Ex. a, b, c, d, e, f*.

13. Sodium was obtained by Sir H. Davy in 1807, by a similar process to that whereby he obtained potassium, using, of course, the hydrate of sodium, NaHO, instead of the hydrate of potassium.

Sodium is now obtained either from the hydrate or carbonate of sodium, by processes similar to those described in § 4 for obtaining potassium.

14. **Soda**, or Protoxide of Sodium, $Na_2O = 62$.

Hydrate of Sodium, $NaHO = 40$.

These substances have similar properties to those possessed by the corresponding compounds of potassium, though in a somewhat feebler degree, and are prepared in a similar way; it will be well, therefore, in order to save repetition, to refer at once to §§ 6, 7, 8, making the necessary changes in nomenclature.

15. Soda, being widely diffused throughout the mineral kingdom, was formerly termed the *mineral* alkali, to distinguish it from potass, which, as being procured from wood ashes, was termed the *vegetable* alkali.

Ex. l.—Try *Ex. g*. with the ashes of burned sea-weed,—first an impure carbonate, and afterwards a hydrate of sodium will be obtained.

16. Soda is powerfully alkaline, neutralizing the strongest acids; it is extremely acrid to the taste, and very caustic. It is employed in the manufacture of glass, and in making *hard* soap.

17. **Lithium**, $Li = 7$. This is a rare metal, somewhat re-

sembling sodium, obtained from a few minerals, and forming the base of an alkali termed **Lithia**, Li_2O. Very little is known about lithium or its compounds; the latter give a carmine color to flame.

18. **Ammonium**, $NH_4=18$. This is a *hypothetical** substance, being, according to some chemists, the metallic base of Ammonia, which is a compound, possessing properties analogous † to those of potass and soda, and exactly replacing them in combination. Ammonium has never been obtained in a separate state, and if it really exists, it differs greatly from the other metals, inasmuch as it is a compound of Hydrogen and Nitrogen, and not a simple body.

19. **Ammonia**, $NH_3=17$. This is a gaseous body, formed by the union of hydrogen and nitrogen. It is colorless, irrespirable, and of a pungent, acrid taste. It possesses powerful alkaline properties, and is lighter than common air (sp. gr. 0.59).

If the gas be *perfectly dry*, it has no action on vegetable colors. (Compare Chap. VII § 14.)

20. Ammonia has a strong affinity for water, which readily absorbs more than 700 times its bulk of the gas, acquiring its odor, taste, and alkaline properties. This solution, NH_3,H_2O, is called "*liquor ammoniæ*," or "*spirits of hartshorn*," and is the common form in which ammonia is sold and used. It has been considered by some as a hydrate of ammonium $(NH_4)HO$ corresponding to the hydrate of potassium, KHO; or the hydrate of sodium, NaHO.

21. The name ammonia is derived from the old name given to chloride of ammonium, NH_4Cl, *sal-ammoniacum*, the Ammoniac salt, from Ammonia, a district in Northern Africa, whence it was obtained.

* *Hypothetical*—supposed, taken for granted.
† *Analogous*—corresponding to, resembling.

It was formerly called "*spirits of hartshorn*," because it was obtained by the dry distillation of *horns*, bones, hoofs, and other animal matter. It has likewise been termed the *volatile* alkali, from its readiness to assume the gaseous form.

22. Ammonia may be readily prepared by heating together in a retort or flask, one part of sal-ammoniac (chloride of ammonium) and two parts of quick-lime; and collecting the gas over mercury, or by displacement, *i. e.* by carrying the delivery tube *upwards* into *dry* inverted jars.

The decomposition may be shewn as follows—

$$2NH_4Cl + CaO = CaCl_2 + H_2O + 2NH_3.$$

Ammonia may also be prepared by the decomposition of dilute nitric acid by either iron or zinc.

Thus—

$$4Zn + 9HNO_3 = 4(Zn2NO_3) + 3H_2O + NH_3.$$

23. A considerable quantity of ammoniacal vapor is given off during the distillation of coal in the process of gas-making, (Chap. XII., § 19). It may likewise be procured, as already stated, by distilling in an iron retort, bones, horns, hoofs, &c.

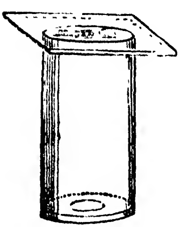


Fig. 63.

In whatever way the gas is prepared, a solution of ammonia may be readily obtained by passing a current of gas through water.

24. Ammonia is one of the most active and valuable elements of manure; it is produced by the putrefaction of all organic substances which contain nitrogen. It exists

in small quantities in atmospheric air, from which, according to Liebig, it is absorbed by the leaves of plants, it enters their roots when dissolved in water, and greatly increases the luxuriance of vegetation.

Ex. n.—Pound together in a mortar small pieces of quick-lime and sal-ammoniac. Though separately they have no scent, they will when thus mingled give off a powerful odor of ammonia, as above explained.

Ex. o.—Pass a current of ammoniacal gas through water colored either with turmeric, reddened litmus, or tincture of red cabbage; the color of the first will be changed to brown, of the second to blue, and of the third, to green.

The experiment may be varied by exposing slips of paper moistened with the colored liquids to the action of the gas; or by adding a few drops of liquid ammonia to the several solutions in test glasses.

Ex. p.—Make a very weak solution of sulphate of copper, and add to it liquid ammonia; hydrated oxide of copper is at first precipitated, but will be instantly redissolved, and form a beautiful blue solution.

Ex. q.—Fill a long, *dry*, glass tube closed at one end with ammoniacal gas; when *quite* full, remove it, keeping it perfectly upright, and immerse it, mouth downwards, in a vessel of water (containing more than sufficient to fill the tube), the gas will be rapidly absorbed by the water, which will rush up the tube with considerable force.

Ex. r.—Fill a bag with ammoniacal gas, and substitute it for the flask B in the adjoining diagram, and pass the gas over binoxide of

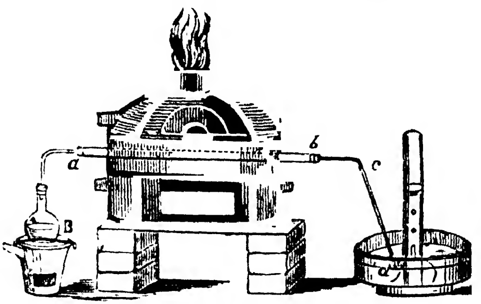


Fig. 64.

manganese heated to redness in the tube *a b*. The ammonia will be decomposed, and its elements uniting with the oxygen from the mineral will form water and nitric oxide gas, which latter may be recog-

nized by the red fumes resulting from its combination with the oxygen of the air—

$$2NH_3 + 7MnO_2 = 7MnO + 3H_2O + 2NO_2.$$

Ex. s.—Perform *Ex. g., Chap. XII.*, using 2 measures of ammoniacal gas and $1\frac{1}{2}$ of oxygen, explosion will ensue. The oxygen will unite with the hydrogen to form water; the nitrogen will be disengaged; some of the oxygen will unite with the nitrogen to form nitric acid, hence the water in the bottle will be *slightly* acid.

Ex. t.—Heat a globule of sodium (or potassium) with a little mercury in a test-tube, they will combine and form an amalgam. When cold, pour over it a strong solution of chloride of ammonium. The mercury will greatly increase in bulk, without losing its lustre, and fill the tube with a light pasty amalgam. "The chlorine and the sodium unite to form common salt, while the mercury combines with the ammonium, *without losing its metallic lustre.* Now there is no instance of mercury or any other metal retaining its metallic properties when combined with a non-metallic substance. The inference is, therefore, that ammonium is a metal."—*Porter.*

EXERCISES ON CHAPTER XX.

1. State the symbol, combining equivalent, and leading physical properties of potassium.
2. What is the most striking chemical property of potassium?
3. When, and by what means was potassium first obtained?
4. How is potassium generally procured?
5. Why is potassium a valuable chemical agent?
6. Name some of the characteristics of pure potass.
7. What is hydrate of potassium?
8. How may hydrate of potassium be procured?
9. Where is potass found in nature? Whence does it obtain its different names?
10. Name some of the properties and uses of potass.
11. State the symbol, combining equivalent, and leading characteristics of sodium.
12. How is sodium distributed in nature?
13. How may sodium be procured?
14. State the properties of soda, and of the hydrate of sodium.
15. Why has soda been termed the *mineral* alkali?
16. Name some of the uses of soda.
17. What is lithium?
18. What is ammonium?
19. Describe the leading properties of ammonia.

20. What is liquid ammonia?
21. By what names is ammonia known, and whence are they derived?
22. How may ammoniacal gas be prepared?
23. How may liquid ammonia be obtained?
24. Of what use is ammonia?

CHAPTER XXI.

METALS OF THE ALKALINE EARTHS.

BARIUM, STRONTIUM, CALCIUM, MAGNESIUM.

Barium.	*Symbol*	Ba.	*Combining*	*weight*	137.
Strontium.	,,	Sr.	,,	,,	87.5.
Calcium.	,,	Ca.	,,	,,	40.
Magnesium.	,,	Mg.	,,	,,	24.

1. **Barium** is a silvery-looking metal, prepared from baryta by a similar process to that described for obtaining potassium. It decomposes water, and oxidizes on exposure to air.

2. It occurs in nature in combination with sulphur and carbonic acid, forming veinstone, which is frequently found in lead mines, and is called by the miners heavy spar, in allusion to its great weight; (βαρυς, *barus*, heavy.)

3. **Baryta**, or Barytes, $BaO = 153$.—This is the oxide of the metal Barium, and when anhydrous, is a grey, porous mass, that slakes, like lime, when moistened with water, forming the hydrate of barium, BaH_2O_2. All the solutions of baryta are more or less poisonous.

4. **Strontium** is a metal similar to barium, but darker in colour, and produced from strontia by a like process; it is found in the lead mines of Strontian, in Argyleshire, whence it derives its name.

5. **Strontia**, $SrO = 103.5$.—This is the oxide of the metal, strontium, and resembles baryta in almost every particular.

Ex. a.—Strontia and its compounds may be instantly known by the brilliant red color which they give to flame; this may be seen by placing a crystal of the nitrate of strontia or some powdered strontia into the flame of a candle or spirit lamp.

6. **Calcium** is similar in appearance and properties to the last two metals; it may be obtained, though with difficulty, from lime, by fusing it with potassium.

7. **Lime**, or the oxide of calcium, $CaO = 28$.—This is an extremely important compound; it is prepared for building and agricultural purposes, by heating chalk or other limestones to redness for some hours in a kiln. By this process the chalk, which is a carbonate of lime, is decomposed, and the carbonic acid is driven off, while the lime remains. When moistened with water it slakes, evolving much heat, and crumbling into a soft powder, which is the hydrate of calcium, CaH_2O_2. This hydrate parts with the water again if exposed to a red heat.

8. The hydrate of calcium is slightly soluble in water, and it is remarkable that the *colder* the water, the *larger* the quantity of lime dissolved. The solution is called lime-water, and is kept as a test for carbonic acid.

9. Lime-water has a disagreeable taste, is acrid and caustic, like solutions of soda and potass. It has a strong alkaline character, turning vegetable blues green, and yellows of a reddish-brown, and restoring the blue color of reddened litmus paper.

10. By exposure to the air, it becomes covered with a film of carbonate of calcium, $CaCO_3$, which is due to the absorption of carbonic acid gas from the atmosphere. The hardening of mortar, which consists of lime and sand, is owing to the same cause.

11. Besides being used for making mortar, lime is

employed in agriculture, to increase the fertility of the soil. Its action in this respect is not well understood, but its effects are most striking on clayey soils and on lands which contain an excess of vegetable matter, as turf or peat.

12. It probably improves the peaty soils by causing the rapid decomposition or decay of the vegetable matter contained in them ; and the clayey soils, by decomposing the insoluble compounds of potass (silicates, &c.) and setting that alkali free to be dissolved by the rain-water and carried into the tissues of plants.

Ex. b.—Place a few small pieces of chalk in a crucible, and subject them to a red heat for an hour or so. This will drive off the carbonic acid, and leave the lime pure.

Baryta and strontia may be obtained by treating their carbonates or nitrates in a similar manner.

Ex. c.—Stir a little lime in some cold water, and allow it to settle, the clear solution is lime-water. Dip a piece of reddened litmus paper into some of it, and the blue color will be restored, showing that it has alkaline properties.

The alkaline properties of baryta and strontia may be shown in a similar manner.

Ex. d.—Boil some of the clear, cold lime-water in a glass flask ; as it becomes warm, very small crystals of lime will be deposited, which will be again dissolved as the liquid cools.

13. **Magnesium** is a soft white metal prepared from the chloride of magnesium, $MgCl_2$, by fusing it with potassium. It burns when heated in the air, producing a brilliant white light, suitable for illumination and for photographic purposes, and forming a white powder well known as magnesia.

14. **Magnesia,** $MgO=40$.—This is the only oxide of the metal, and is prepared by raising to a white heat in an earthen crucible the common magnesia of the shops, which is a carbonate. As in the case of lime, the carbonic acid is driven off by heat, and there remains, pure, or as it is sometimes called, calcined magnesia. It is a soft white powder, attracting moisture and carbonic acid from the atmosphere. Its alkaline reaction is not so easily shown

as that of the previous oxides, out it completely neutralises the most powerful acids.

EXERCISES ON CHAPTER XXI.

1. Give the names, symbols and equivalents of the metals of the alkaline earths. State the properties of barium.
2. Name some natural combinations of barium.
3. Describe the nature and properties of baryta.
4. What is strontium?
5. Give the symbol, &c. for strontia.
6. How may calcium be obtained?
7. How is lime procured?
8. What is lime-water?
9. Describe the properties of lime-water.
10. Why does mortar harden on exposure to the air?
11. On what kind of soils is lime most useful?
12. How does lime act in promoting the fertility of the soil?
13. Describe the metal magnesium.
14. What are the properties of magnesia?

CHAPTER XXII.

METALS OF THE EARTHS PROPER.

ALUMINUM, OR ALUMINIUM, &c.

Aluminium. *Symbol* Al. *Combining weight* 27.5.

1. **Aluminium** is prepared by decomposing the chloride of Aluminium, Al_2Cl_6, by means of metallic sodium. It is a white metal, having the appearance of tin, but much lighter. It does not readily tarnish in the air, and may indeed be poured in a melted state from one vessel to another without oxidation (sp. gr. 2.6).

2. It has recently been prepared in somewhat large quantities, with a view to its employment in the arts. A cheap method of obtaining potassium or sodium, would no doubt lead to the more general use of this metal.

3. It is highly sonorous, and therefore suited for the

manufacture of bells. On account of its lightness and tenacity the French government have proposed using it for helmets and cuirasses. Its power of receiving and retaining a brilliant polish, combined with its lightness, would cause it to be used for many purposes in which silver is now employed.

4. **Alumina**, or the sesquioxide of aluminium, $Al_2O_3 =$ **103.**—This substance is one of the most abundant productions of nature. It is found in every region of the globe, and in rocks of all ages. The different kinds of clay of which bricks, pipes, and earthenware are made, consist of silicate of aluminium in a greater or less degree of purity.

5. This earth generally appears in rude shapeless masses, but sometimes it is found beautifully crystallized. The ruby and the sapphire are composed of nearly pure alumina with a little coloring oxide.

6. Alumina is mostly found combined with silica and potash, as in the case of the mineral termed felspar, which forms the chief constituent of granite, porphyry, and other unstratified rocks, and may be considered a double silicate of potassium and aluminium, $K_4SiO_4 + Al_43SiO_4$.

7 The gradual crumbling down of hard granite rocks, is owing to the decomposition of the felspar and the removal of the silicate of potassium; the result is clay, in a greater or less degree of purity. The white porcelain clay of Devonshire and Cornwall is thus produced from the fine white granite of those districts.

Clays are often colored, owing to the presence of oxide of iron or other substances. Clays which contain chalk or carbonate of calcium are termed *marls*, and may be known by their effervescing with acids.

8. Pure Alumina has neither taste nor smell, is of a white color, and is quite insoluble in water, which, however, it readily absorbs. It has a strong affinity for various coloring matters; hence, it is of great importance in dyeing: the pigments called lake colors, such as crimson

lake, madder lake, &c., are coloring matters combined with alumina.

Ex. a.—Boil cochineal in water until its color is extracted; in this infusion dissolve some alum (which is a sulphate of alumina and potass), and then add ammonia. This will take away the sulphuric acid of the alum, and precipitate the alumina, which will carry down with it the coloring matter of the cochineal, and leave the liquid clear and colorless. The precipitate is carmine. In the same way numerous other colors may be made.

9. The other metals of this class occur only in very small quantities, or in extremely rare minerals, and have therefore been passed over. Their names are **Cerium, Didymium, Erbium, Glucinium, Lanthanum, Norium, Terbium, Thorium, Yttrium,** and **Zirconium.**

EXERCISES ON CHAPTER XXII.

1. State the symbol, equivalent, and properties of aluminium.
2. Why is not aluminium in common use?
3. For what purposes might aluminium be employed?
4. What is alumina?
5. Name some natural forms of alumina.
6. What is felspar?
7. How is porcelain clay produced? What is marl?
8. What are the properties of pure alumina?
9. Why are the other metals of this class comparatively unimportant?

CHAPTER XXIII.

METALS WHOSE OXIDES FORM POWERFUL BASES.

Group A—*Not precipitable from* acid *solutions by sulphuretted hydrogen.*

MANGANESE, IRON, CHROMIUM, NICKEL, COBALT, ZINC.

Manganese,	*Symbol*	M,	*Combining*	*weight*	55.
Iron,	,,	Fe,	,,	,,	56.
Chromium,	,,	Cr,	,,	,,	52.5.
Nickel,	,,	Ni,	,,	,,	59.
Cobalt,	,,	Co,	,,	,,	59.
Zinc,	,,	Zn,	,,	,,	65.2

1. **Manganese** is a grey metal, something like cast-iron,

very hard and brittle. It fuses with difficulty, and when pure oxidises so rapidly, that, like potassium, it must be kept in naphtha. It is found in nature chiefly in the form of the black oxide or peroxide, MnO_2, and as a constituent of many other minerals.

2. The black oxide is employed in the preparation of oxygen and chlorine, as already stated; it is also used to give a violet tint to glass, and in the production of artificial amethysts. There are several other oxides of this metal, some having acid properties, but they are of comparatively little importance.

Ex. a.—Mix together in a mortar equal parts of salt-petre (nitrate of potassium) and black oxide of manganese; put this mixture into a crucible large enough to hold three times the quantity, and expose it to a red heat for half-an-hour, or as long as any gas is given off. When cold, add a little water to the mass, a green solution of manganate of potassium will be obtained, which may be poured off, evaporated, and crystallized. It has been called the *mineral chameleon* from the curious changes it undergoes.

Ex. b.—Put some of the clear solution formed in the preceding experiment, or 10 or 12 grains of the crystals into a tall glass, and add gradually some water; the green color soon begins to change through light green, blue, purple, to a crimson-red. If a few drops of sulphuric acid be added to the green solution, the change to red is instantaneous. These alterations of color arise from the salt attracting oxygen.

Ex. c.—Put into either the green or red solution, a piece of wood or paper, or other organic substance, and the color will be removed, so easily is the acid decomposed.

Ex. d.—Put some of the crystals into two glasses, and add to the one *hot*, and to the other, *cold*, water. The hot solution will be of a beautiful green color, and the cold one of a deep purple.

Ex. e.—Pound some flint-glass in a mortar, and add to it a very little black oxide of manganese; fuse the mixture by the blowpipe or in a crucible. The glass will become of a beautiful amethyst color.

3. **Iron** is a most important metal, it occurs native to a small extent. When pure, it is of a bluish-white color, has a metallic lustre, is very soft and tough, exceedingly malleable, and highly tenacious.

4. The ordinary source of this metal is the clay iron-stone which is found in thin beds or nodules, associated with coal. The iron-stone consists of the oxide of iron, carbonic acid, silica, clay, lime, magnesia, and other substances. It is first broken to pieces and heated in a furnace to expel the water and carbonic acid, and is then ready for reduction to the metallic state. The furnace in which this process of reduction or smelting is carried on, is usually of large dimensions, sometimes fifty or sixty feet high, the interior being lined with fire-bricks. The bottom of the furnace is closed, and a blast of air is introduced by means of pipes. The roasted iron ore, which

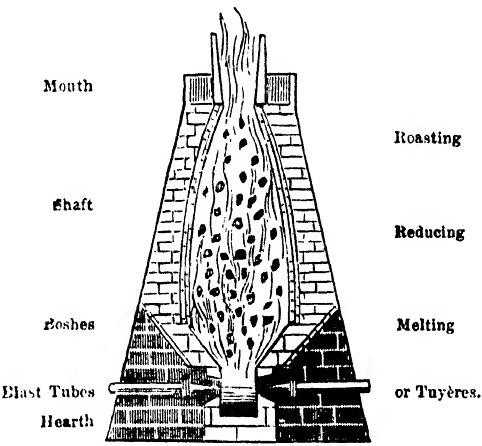


Fig. 65.

contains on an average about 30 per cent. of metal, is introduced at the top, mixed with coal and limestone. It is worthy of observation, that the materials necessary for the reduction of the iron are all found in the same locality. When once the furnace is in operation, the smelting continues for years without intermission, fresh materials being

from time to time introduced at the top. It is thought that the iron is reduced to its metallic state in the central parts of the furnace, where the temperature is very high. As the mass sinks down in the furnace to the part termed the *boshes*, it becomes more strongly heated; the iron unites with the white hot carbon, forming carbide or carburet of iron, while the silica and alumina unite with the lime of the limestone to form a glassy slag. The iron and slag now reach the bottom or *hearth* in a fluid state, and arrange themselves according to their weight; the molten iron at the bottom, and the slag floating on its surface. These are then drawn off through different openings at the bottom of the furnace, and the iron being run into moulds of sand, constitutes the cast-iron or pig-iron of commerce.

5. Formerly these furnaces were heated with coke, but by the introduction of hot air through the *blast pipes*, raw coal may be used instead. The air is raised to a temperature of 600° or 700° by being made to traverse a series of red hot iron pipes, it is then forced into the furnace by a blower, worked by a steam engine. An immense saving of fuel is effected by this arrangement, which is termed the hot blast.

6. The cast-iron or pig-iron of commerce is converted into malleable wrought-iron by a process called *puddling*. This process is conducted in a reverberating furnace. The cast-iron is laid on the bed of this reverberating furnace, a current of burning gases proceeds from the fuel, which is reflected to the bed of the furnace by the arched roof. The heat soon

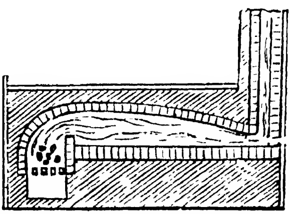


Fig. 66.

melts the metal, which becomes covered with a crust of the oxide of iron. The workman, with an iron rod like a boat

oar, diligently stirs the melted mass, so as thoroughly to mix the oxide with the metal, and bring every part of the molten mass to the surface. A little water is occasionally thrown on the metal, to promote a more rapid absorption of oxygen. The object of the process is to convert the carbon of the cast-iron into carbonic oxide, CO, and jets of this gas may be seen burning on the surface, with its characteristic blue flame. The mass gradually loses its fluidity, and becomes pasty and granular. The heat of the furnace is now quickly raised, the workmen gather the particles into large balls, which are removed from the furnace, and placed under a large hammer—by which most of the fluid impurities are squeezed or hammered out. The iron is then heated a second time, and passed between large iron rollers, furnished with grooves; by which process it is converted into the wrought bar-iron of commerce.

7. When very superior iron is required, the bars thus made are cut into a number of pieces, which are bound together and again raised to a white heat, hammered and rolled into a single bar. This process is sometimes repeated two or three times, and is called *fagoting*. Wood charcoal is the best fuel for the reduction of iron, but the expense is too great to admit of its use.

8. Steel, which is a compound of carbon and iron, is made from cast-iron, by burning out half its carbon, or from wrought-iron by restoring half the carbon of which it was deprived in the *puddling* furnace. The latter is the process generally adopted, and is termed *cementation*. It consists in burying bars of the best malleable iron in boxes containing charcoal powder, and strongly heating them for several days. Under these circumstances, the iron takes up about one 150th part of its weight of carbon, becoming harder and more elastic, but less malleable than before.

9. The degree of hardness depends on the rate of cooling. Steel, if heated to redness and cooled slowly, becomes almost as soft and malleable as wrought-iron. This process is called *annealing*.

10. If cooled very suddenly, it becomes as hard and brittle as cast-iron; in this state it is unfit for most uses, and is therefore subjected to another process termed *tempering*. The sort of temper attained, depends upon the degree of heat imparted; for any portion of its original softness and malleability may be restored, by reheating and slow cooling.

11. The proper degree of heat may be ascertained by the color assumed by the heated steel. Tools for working metals are heated to a pale yellow; planes and knives to a darker yellow; chisels and hatchets to a purplish yellow; springs to a full purple. In each case the metal is allowed to cool slowly.

12. The colors are owing to the different degrees of oxidation to which the steel is subjected. They may all be seen by heating a steel knitting needle in the flame of a lamp. Where the heat is greatest, the needle becomes blue or purple, shading off into pale yellow on either side.

13. The toughness of wrought-iron depends chiefly on its fibrous character, but this toughness slowly disappears when the iron is subject to continued hammering or jarring, in consequence, it is believed, of the development of a crystalline structure, similar to that seen in cast-iron when broken. Accidents often happen through the breaking of axles and shafts, in which this change of structure has taken place.

14. Wrought-iron becomes soft at a white heat, and in this state two pieces may be united into one by hammering. This property, which adds greatly to the usefulness of iron, belongs to no other metal except platinum. The process is termed *welding*.

15. Another distinguishing property of iron is its susceptibility of magnetism. The loadstone itself, in which this power was first noticed, is an iron ore.

Magnetism may be communicated to iron by contact with a loadstone or an artificial magnet, but immediately dis-

appears from *soft* iron when the magnet is removed; steel, on the contrary, retains its magnetic power, and does not lose it until heated to redness.

16. In air absolutely dry, and in water entirely free from air, iron undergoes no change; but when both air and moisture act upon it together, it speedily rusts or oxidises.

17. **Protoxide of Iron,** $FeO=72$:—This is a very powerful *base,* completely neutralizing acids. It is almost unknown in a separate state, from its great readiness to absorb oxygen and pass into the sesqui-oxide. The soluble compounds which it makes with acids have generally a pale green colour, and a disagreeable metallic taste.

18. **Peroxide of Iron.** $Fe_2O_3=160$:—This body, which is a sesqui-oxide of iron, and is nearly the same as rust of iron, occurs native, and is known in mineralogy as *specular iron ore,* and as *red* or *brown hæmatite.* It combines with acids, forming compounds usually of a reddish color.

19. **Black Oxide of Iron.**—When iron is heated to redness in the open air, it rapidly absorbs oxygen, and is converted into black scales. Such scales may be seen flying off from a mass of red-hot iron as the blacksmith hammers it on his anvil. These scales, which have been termed by some, the black oxide of iron, and by others, the protoxide, are rather to be considered as a mixture of protoxide and peroxide, $FeO+Fe_2O_3$. As a natural product, it is known as the *magnetic oxide* of iron, or loadstone.

Ex. f.—Wind a piece of fine iron or steel wire round a lead pencil so that it may form a coil, fasten a bit of thread to the end, and dip it in sulphur. Inflame the sulphur, and immerse it while burning, into a jar of oxygen, the wire will immediately begin to burn, and fly off in brilliant sparks. These sparks, if examined when cold, will be found very different from the metal from which they were formed, being brittle and without metallic lustre; they are the black oxide of iron. The inside of the jar will be seen covered with a red powder, this is the peroxide of the metal. When performing this experiment, some water or

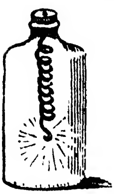


Fig. 67.

wet sand should be placed at the bottom of the jar, in order to catch the melted particles, which are so hot that they would otherwise crack the jar; the jar, too, should be rather wide, lest the sparks should strike its side and produce the same result.

Ex. g.—To distinguish iron from steel. Let a drop of diluted nitric acid fall upon the metal; and after a few minutes, wash it off with water. If the metal be steel, a black spot will be left on it; if it be iron, a whitish grey spot will remain. The acid dissolves the iron in both cases, but the carbon of the steel remains undissolved, and produces the blackness.

20. **Chromium** derives its name from the Greek word *χρῶμα*, *chroma*, color, on account of the beautiful color of many of its compounds, which are largely used both in painting and dyeing.

It is a hard, greyish-white, brittle metal, and, like manganese and pure iron, very difficult to be fused.

21. Chromium is found chiefly in *chrome iron-stone*, a mineral somewhat abundant in the Shetland Isles and elsewhere, and consisting of oxide of iron and sesqui-oxide of chromium, $FeO + Cr_2O_3$. It forms four compounds with oxygen, corresponding to the compounds of oxygen and iron.

22. **Protoxide of Chromium**, $CrO = 68 5$:—This is a powerful base, forming pale blue compounds with acids, and rapidly absorbing oxygen.

23. **Sesqui-oxide of Chromium**, $Cr_2O_3 = 153$:—This substance, which is of a fine green colour, and much used in painting on glass, porcelain, and enamel, may be obtained in the anhydrous state by heating bichromate of potassium, $K_2O,2CrO_3$, or rather $K_2Cr_2O_7$, to redness. One half the acid is decomposed, oxygen gas being liberated, and sesqui-oxide of chromium remaining mixed with chromate of potass: the latter may be dissolved out by water, and the sesqui-oxide washed and dried.

$$2(K_2Cr_2O_7) = 2(K_2CrO_4) + 3O + Cr_2O_3.$$

24. **Chromic acid**, $H_2CrO_4 = 118.5$.—When oxide of chromium is heated with an alkali in the air, oxygen is absorbed, and a chromate of the alkali formed.

Chromic acid may be obtained nearly pure by mixing a cold saturated solution of bichromate of potassium with half as much again by measure of sulphuric acid.

$$K_2Cr_2O_7 + 2H_2SO_4 + H_2O = 2KHSO_4 + 2H_2CrO_4.$$

Chromic acid acts upon and destroys vegetable colours, and is hence used to bleach tallow and palm oil previous to their manufacture into soap and candles.

Its compounds are of considerable importance, and will be noticed under the head of chromates.

25. **Nickel** was discovered in 1751. It is found (usually combined with arsenic) in the Harz mountains, and some other places, and also in meteoric stones. The old German miners gave it the name of *kupfer-nikel*, thinking it a kind of *false copper* ore.

26. Nickel is white, brilliant, ductile, and malleable, and strongly magnetic. Its principal use in the arts is for the preparation of the alloy termed German silver, which is a mixture of 100 parts of copper, 60 of zinc, and 40 of nickel. This alloy is very malleable, and will take a high polish.

27. **Cobalt** is a metal of a reddish grey color, hard, brittle, and fused with difficulty; it is also strongly magnetic. It greatly resembles nickel, with which metal and arsenic it is generally associated.

28. **Oxide of Cobalt**, $CoO = 75$:—This substance imparts a magnificent blue color to glass: the pigment called *smalt* consists of glass so colored and ground to powder. Zaffre and cobalt-ultramarine are likewise preparations of cobalt.

Ex. h.—Perform experiment *e* as described under manganese, using oxide of cobalt instead of that metal, and the result will be a beautiful blue glass.

29. **Zinc** is a somewhat abundant metal, of a bluish white color; and brittle at ordinary temperatures. At a heat, however, somewhat above that of boiling water (250°—300°), it is malleable and ductile, and hence it is now used for numerous purposes instead of lead. At a higher temperature it again becomes brittle, so that at 400° it may be reduced to powder. At 773° it fuses, and crystallizes if cooled slowly; at a red heat it boils and goes off in vapor.

30. The most valuable of the ores of zinc are the native carbonate or *lapis calaminaris*, and the sulphuret or *blende*; they are found associated with lead ores in Great Britain and other parts of the world.

31. Zinc is obtained by placing the carbonate or sulphuret of the metal (the latter previously roasted to expel the sulphur) into a large earthen retort, with one-fifth its weight of charcoal; this mixture being raised to a red heat, the ore is decomposed, the charcoal combining with the oxygen, escapes into the air, while the reduced metal distils over as a vapor, and is condensed into the metallic state in a vessel of water, into which the tube of the retort is immersed.

32. Zinc is now used for many articles formerly made of copper, lead, or iron, as it is cheaper than copper, harder and lighter than lead, and less liable than iron to be injured by air and water. It is sometimes used as a coating to protect iron chains, &c., from rusting. This coating is produced by plunging the iron, previously made perfectly clean, into melted zinc, which forms an alloy on its surface. Iron so covered has been called, but without reason, *Galvanized* Iron.

Zinc is extensively employed in the arts, in combination with copper, with which it forms the well-known alloys, brass, mosaic gold, pinchbeck, &c.

33. **Oxide of Zinc**, $ZnO = 81$: — Zinc soon tarnishes when exposed to moist air, becoming covered with a thin film of oxide, that resists further change.

When the metal is raised to a red heat, in contact with the air, it takes fire and burns with a whitish green flame, being converted into oxide of zinc, which escapes in dense white fumes, that finally settle into a wool-like substance, called by the old chemists *Philosopher's wool.*

34. The oxide of zinc may also be prepared by heating the carbonate of the metal to redness, or by adding caustic potass to a solution of any salt of zinc, such as the sulphate; in the latter case, a hydrate of the oxide, ZnH_2O_2, will be precipitated, which is soluble in an excess of the alkali.

35. Oxide of zinc is a white powder, inodorous, tasteless, insoluble in water, and infusible by heat. At a low red heat it becomes yellow, but recovers its whiteness on cooling. It has been proposed to use oxide of zinc as a paint instead of white lead (carbonate of lead), because it is prepared with less danger to the health of the workmen, and retains its color better when exposed to a smoky atmosphere.

Ex. i.—Melt some pieces of zinc in an iron ladle: drop the melted metal, little by little, into a pail of water: the hardened drops are called *granulated zinc*, and may be removed and kept for making hydrogen and for other purposes. When all the metal has been poured away, turn the dross out of the ladle into the hottest part of the fire. The metallic particles contained in it will burn, as above described, with a beautiful flame.

Ex. k.—Put some zinc filings or turnings into a deflagrating spoon, together with a small piece of phosphorus, light the latter and plunge the spoon into a jar of oxygen; the metal will burn, and give rise to dense white fumes of oxide of zinc, ZnO.

Ex. l.—Hold a leaf of *white* Dutch metal (which is leaf zinc) in the flame of a spirit lamp; it will burn with a vivid flame.

Ex. m.—Melt three or four pounds of zinc; pour it into a crucible which has a small hole in the bottom, stopped with a plug of wood. When the surface of the melted metal begins to congeal, pull out the plug by a string or wire, so that the still fluid metal may escape into a vessel placed beneath to receive it. Upon afterwards breaking the

crust in the crucible, the inner surface will be found beautifully crystallized, the crystals being of a prismatic form and quite brilliant. (To preserve the brilliancy, they should be immediately varnished.)

EXERCISES ON CHAPTER XXIII.

1. Give the symbol, equivalent, and leading properties of manganese.
2. For what is the black oxide of manganese used?
3. Give the symbol, &c. of iron.
4. Describe the process of smelting iron ores.
5. What is meant by the hot blast?
6. Describe the process for rendering cast-iron malleable.
7. What is meant by *fagoting?*
8. How is steel made?
9. What is annealing?
10. What is meant by tempering steel?
11. How is the proper heat for tempering steel ascertained?
12. To what are the colors on heated steel owing?
13. On what does the toughness of wrought-iron depend?
14. What is meant by welding?
15. What is load-stone? How does iron act with regard to magnetic influences?
16. How is iron affected by dry air?
17. Describe the protoxide of iron.
18. State the properties of peroxide of iron.
19. What is the black oxide of iron?
20. What is chromium? and whence does it derive its name?
21. Where is chromium chiefly found?
22. What is protoxide of chromium?
23. How is the sesqui-oxide of chromium prepared, and for what is it used?
24. Describe the preparation and properties of chromic acid.
25. When and where was nickel discovered? Whence does it derive its name?
26. State the chief properties and uses of nickel.
27. State the symbol and characteristics of cobalt.
28. What are the properties of oxide of cobalt?
29. Give the symbol, equivalent, and leading properties of zinc.
30. Which are the principal ores of zinc?
31. How is zinc obtained?
32. For what is zinc used?
33. How is oxide of zinc prepared?
34. How is the hydrated oxide of zinc obtained?
35. For what is oxide of zinc used?

CHAPTER XXIV

METALS WHOSE OXIDES FORM POWERFUL BASES—*continued.*

Group B—*Precipitable from their solutions by sulphuretted hydrogen.*

COPPER, LEAD, BISMUTH, CADMIUM.

Copper.	*Symbol* Cu ;	*Combining weight*		63.5.
Lead.	,, Pb ;	,,	,,	207.
Bismuth.	,, Bi ;	,,	,,	210.
Cadmium.	,, Cd ;	,,	,,	112.

1. **Copper** appears to have been one of the earliest metals employed by mankind, having been used for domestic utensils and weapons of war before the discovery of malleable iron.

The ancients obtained very considerable quantities from the island of Cyprus (Κυπρος, *cupros*), and this is supposed to be the origin of the name copper (*cuprum*).

2. Copper is found native or in the metallic state, and also in various forms of combination.

Native copper is found in abundance on the southern shores of Lake Superior.

The most important of its ores are:—the *yellow copper ore* or *copper pyrites*, which is a double sulphuret of copper and iron, $CuS + FeS$;—the *red oxide* or suboxide of copper, frequently termed *ruby copper*, Cu_2O;—and *sub-carbonate of copper*, or *malachite*, $CuCO_3,CuH_2O_2$.

3. Large quantities of the last two mentioned ores have been of late obtained from South Australia. The mines of Devonshire and Cornwall yield chiefly the *yellow copper*

ore, which is shipped to Swansea, in South Wales, to smelted there, on account of the greater abundance of f

"By this arrangement the vessels not only carry the smaller quan of material to the greater, but load back with coal for the use of mines."

4. The smelting process is as follows.—The ore is f *calcined*, or roasted, in a reverberatory furnace, whereby the sulphur, &c. is burnt out, going off in the form of sulphurous and sulphuric acids, while the copper and iron are both oxidized. The ore thus calcined is strongly heated with charcoal and sand. Metallic copper, Cu, and carbonic oxide, CO, are

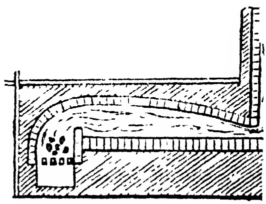


Fig. 68.

formed from the oxide of copper, CuO, and the charcc C; while the sand, SiO_2, unites with the oxide of iron, F forming a silicate of iron, Fe_2SiO_4, which floats slag on the surface of the molten metal.

The roasting and melting processes must be alternately repe several times before pure copper can be obtained.

The formation of the slag with the oxide of iron, depends on fact, that oxide of iron is not so easily reduced to the metallic stat oxide of copper.

5. Copper is a tough, sonorous metal, of a red co and nearly nine times as heavy as water. It is mallea both when hot and when cold, and may be reduced to t leaves; it is also very ductile, and may be drawn into wire. It is, moreover, the most tenacious metal nex iron; a wire $\frac{1}{10}$th of an inch in diameter, being able support 302lbs without breaking.

Copper is softened by heating and quenching in wat the reverse of the effect produced on steel by such tre ment.

6. In dry air copper oxidizes but slowly, gradually

quiring a brown tarnish, which consists of a merely superficial film of sub-oxide of copper, Cu_2O.

In moist air copper oxidizes much more rapidly, soon becoming covered with a green crust which is a carbonate of copper.

7. **Protoxide or Black oxide of Copper, $CuO = 79.5$:** This substance, which is the base of the ordinary salts of copper, is prepared by heating copper to redness, with free access of air, or by decomposing the nitrate of copper by a red heat. It is a dense, black powder, insoluble in water and the fixed alkalies, but soluble in ammonia and acids.

Heated with charcoal, or in contact with organic matter, it is reduced either to metallic copper or to the sub-oxide.

8. **Suboxide or Red oxide of Copper, $Cu_2O = 143$:**— This compound is frequently found in Cornwall in the form of beautiful transparent crystals of a fine red color. It forms colorless compounds with acids. The sub-oxide of copper communicates a red tint to glass, while that given by the protoxide is either green or blue.

9. Sheet, or rolled copper, is used for sheathing ships, roofing buildings, making boilers, kettles, and other utensils, for which iron, on account of its more rapid oxidation, would be less suitable.

Vegetable acids dissolve copper when cold, but not when hot; hence copper vessels, if kept *clean*, are not dangerous for cooking, provided what is boiled in such vessels be not allowed to cool in them, but be instantly poured out.

Sauces containing vinegar, preserved fruits, jellies, &c., should not be suffered to *remain* in copper vessels, as all the compounds of copper are poisonous.

In case of poisoning by copper, the best antidotes are, first, an emetic, and afterwards, milk, wheat-flour, white of eggs, iron filings in gum water, sugar or any syrup. Vinegar must be specially avoided.

Ex. a.—Hold a bright copper coin, by a pair of pincers, in the flame of a lamp: notice the changes of color, owing to the various degrees of oxidation. If kept in the point of the flame for some time, it becomes covered with black oxide, on suddenly quenching it in water, this oxide

falls off in black scales, shewing the reddish brown color of the sub-oxide beneath.

Ex. b.—Immerse a piece of clean iron or steel, as the blade of a knife, in a solution of sulphate of copper (blue vitriol), the surface will be speedily covered with a thin film of metallic copper. By this means the presence of copper in preserves, pickles, ink, tea, &c., may be readily detected

Ex. c.—Repeat *Ex. g.*, chap. xiv.

Ex. d.—Add a solution of caustic potass to a solution of sulphate (or nitrate) of copper, a blue precipitate of the hydrated oxide of copper, CuH_2O_2, will be thrown down; this, when heated, parts with the water, and is converted into the black protoxide, CuO.

$$\text{1st. } CuSO_4 + 2KHO = K_2SO_4 + CuH_2O_2.$$

$$\text{2nd. } CuH_2O_2 = H_2O + CuO.$$

10. Copper unites with several other metals to form alloys, the most important of which are the following :—

Brass, which consists of different proportions of copper and zinc.

Dutch Metal, which is a kind of brass beaten into thin leaves, and used instead of leaf-gold in imitation gilding and bronzing.

Bronze, Bell Metal, and **Gun Metal**, which are compounds of copper and tin in different proportions, with sometimes a little zinc.

Mosaic gold, which is a kind of brass, consisting of 3 parts copper and 1 zinc, with a little tin.

German silver, which is composed of 2 parts copper, 1 nickel, and 1 zinc.

11. **Lead** is a soft, flexible, bluish-grey metal, easily scratched by the nail, and leaving a black streak upon paper. Its specific gravity is more than eleven times that of water (11.4). It is malleable and ductile, but not very tenacious, since a wire $\frac{1}{10}$th of an inch in diameter breaks with a weight of less than 30lbs.

12. At common temperatures and in dry air, lead tar-

nishes slowly, but in moist air it soon becomes coated with a film of grey matter thought by some to be a suboxide, Pb_2O.

13. Lead fuses at about 612°, and if slowly cooled may be obtained in octohedral* crystals. It is not volatile, for *in close vessels*, it can be heated to whiteness, without subliming.

If the metal be melted in contact with the air, it absorbs oxygen rapidly, and the surface becomes covered with *dross*, a mixture of metallic lead and the protoxide, PbO.

14. Lead occurs in nature under various mineral forms: of these, the most valuable and the most abundantly diffused is sulphuret of lead, or *galena*, PbS, which is found mostly in the older secondary rocks, and is the chief, if not the only source, of the lead of commerce. Galena often contains silver, and sometimes in such quantity as to render it worth separating.

15. Metallic lead is obtained from galena, by exposing the crushed ore to a dull red heat in a reverberatory furnace, whereby the sulphuret, PbS, is oxidized by the heated air and converted partly into sulphate of lead, ($PbSO_4$), and partly into oxide, (PbO), sulphurous acid gas, SO_2, being given off. The temperature is then raised, whereupon the recently formed sulphate and oxide react upon the remaining sulphuret, producing sulphurous acid gas, SO_2, and metallic lead, Pb.

This decomposition may be shewn as follows:—

$$\text{1st stage, } 3PbS + 10O = PbSO_4 + 2PbO + 2SO_2.$$

$$\text{2nd ,, } \begin{cases} PbSO_4 + PbS = 2Pb + 2SO_2. \\ 2PbO + PbS = 3Pb + SO_2. \end{cases}$$

16. The *scoriæ* (or *slag*) which float on the surface of the melted metal, consist of the earthy matter of the ore mixed with a considerable quantity of lead. They are

* *Octohedral*—eight-surfaced. A solid enclosed by eight triangular surfaces is termed an *octohedron*, plural, *octohedra*.

therefore, removed to another furnace and re-melted, whereby a further portion of less pure metal is obtained.

17. **Protoxide of Lead,** $PbO = 223$:—This substance is obtained by exposing melted lead to a current of air, or by heating carbonate of lead (*white lead*) to redness, whereby the carbonic acid is driven off.

If carefully prepared, and not allowed to fuse, it is of a pale yellow color, and is known under the name of *Massicot.*

18. When *moderately heated,* protoxide of lead becomes of a deep red color, but recovers its original yellow tint on cooling.

At a strong *red heat* it fuses, and cools in the form of deep brick-red crystalline scales, which are known by the name of *Litharge.* This substance when boiled with oil gives it a drying quality, and is therefore much used by painters.

19. **Deutoxide of Lead, or Red Lead,** Pb_3O_4, or $2PbO + PbO_2$:—This substance is made by heating the oxide of lead (*massicot*) in a current of air, so as to oxidize without fusing it. It is a brilliantly colored and extremely heavy powder, much used as a cheap red paint.

20. **Peroxide of Lead,** $PbO_2 = 239$: — This compound may be formed by the action of nitric acid on red lead. The red lead is decomposed into protoxide and peroxide; the former unites with the acid, while the latter remains as an insoluble puce-colored powder.

$$Pb_3O_4 + 4HNO_3 = 2(Pb2NO_3) + 2H_2O + PbO_2.$$

21. Lead placed in distilled water, *free from air,* remains unchanged; but if air be present, the lead is gradually oxidized, and the protoxide so formed, being dissolved by the water, renders the latter poisonous. By the action of the carbonic acid absorbed from the atmosphere by the water, the oxide of lead, so dissolved in it, is gradually converted into the carbonate, and precipitated.

The purer the water, the more certain the poisonous action, hence rain-water, or very soft waters, intended for drinking or for cooking purposes, should not flow through leaden pipes, or be kept in leaden cisterns.

22. A small proportion of sulphuric acid or of some soluble sulphate or chloride in the water, entirely prevents the corrosion of the lead. Hence spring water, well water, or river water, which generally contains such bodies, may be safely kept in leaden cisterns, because the sulphuric acid or the chlorine forms an insoluble compound which quickly encrusts the surfaces of the cistern and prevents any further action on the metal.

No water should be used for cooking purposes out of *new* leaden cisterns, but time must be given, and the water left unchanged in order that the encrustation may be formed. This may be hastened, if necessary, by pouring a few drops of sulphuric acid into the water.

Leaden cisterns should have *wooden*, not leaden covers, for the vapor condensed on the inside of the leaden cover, being pure water, would become contaminated with lead, as stated above.

Ex. e.—Place some shavings of lead on ignited charcoal, and throw a stream of oxygen gas upon them from a bag, or bladder, the metal will burn with a beautiful blue flame.

Ex. f.—Rub some peroxide of lead in a mortar with one-fifth its weight of sulphur. If the mortar be hot, the two bodies will combine with great rapidity and burst into flame.

Ex. g.—Dissolve half-an-ounce of sugar of lead (acetate of lead) in water in a large, clear glass bottle; add a few drops of nitric acid or vinegar, to remove any cloudiness, and hang a piece of zinc in the solution by a string fastened to the cork of the bottle. The zinc will soon appear covered with particles of pure lead. This arises from the greater affinity of acetic acid for zinc than for lead. The zinc gradually displaces the lead from the acetate of lead, and the solution becomes acetate of zinc.

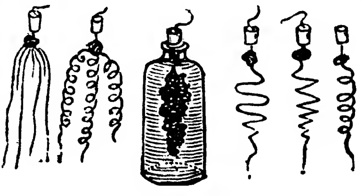


Fig. 69.

If a piece of brass wire be twisted round the zinc as in the diagram, a still more beautiful effect is produced, the lead branching out in leaf-like scales, whence the experiment has been named the lead-tree. The use of the brass

wire introduces a new and additional cause of decomposition, namely, galvanic or electrical action.

23. **Bismuth** is a metal generally found native, but sometimes in union with arsenic, sulphur, or oxygen. It is somewhat abundant in Saxony, Bohemia, and Transylvania; it is also met with in Cornwall and Cumberland. The metal is obtained from the rock which contains it, by simple heating in inclined tubes, for at a moderate temperature the bismuth fuses and runs down into vessels place to receive it.

24. Bismuth has a reddish-white color, and is so brittle that it may be readily reduced to powder; it fuses at 497°, at a higher temperature it volatilizes and may be distilled in close vessels. If a considerable quantity of the metal be melted, and allowed to cool gradually, it crystallizes in beautiful and well-defined cubes.

Ex. h.— Perform *Ex. a*, Chap. xiv., using two or three pounds of bismuth instead of sulphur, the result will be the formation of crystals arranged as in the diagram, and presenting a fine example of metallic crystallization

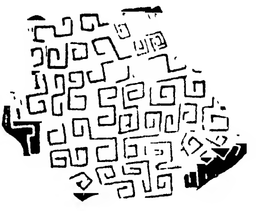


Fig. 70.

The chief use of bismuth is for the purpose of making fusible alloys, stereotype plates, queen's metal, &c., and as an ingredient in solders.

See Experiments *b*, and *c*. Chap. xviii.

25. **Oxide of Bismuth**, $Bi_2O_3 = 468$:—Bismuth oxidizes but slowly in the air at common temperatures, but when heated to whiteness it burns with a bluish flame, and is converted into a yellow fusible powder, which is the protoxide of the metal.

Oxide of bismuth may likewise be obtained by dissolving bismuth in nitric acid, so as to form nitrate of bismuth, $Bi3NO_3 + 5H_2O$; on diluting the solution with water, decomposition takes place, and a subnitrate of bismuth, $2BiNO_4 + H_2O$, is precipitated as a yellowish-white powder; if this be heated to dull redness, the nitric acid is expelled, and oxide of bismuth remains.

Ex. i.—When heated on charcoal before the blow-pipe, bismuth soon fuses and burns, coating the charcoal with a yellow oxide. If the globule of melted metal, be allowed to fall on the floor or into a pasteboard box, it will immediately divide into numerous smaller globules, radiating in all directions, and each leaving a distinct yellow mark behind it.

26. **Cadmium** is a white, malleable, and rather hard metal, discovered by Stromeyer in 1817, and found in small quantities associated with ores of zinc. It fuses at 442°, and is volatile at a low red heat. It tarnishes but little in the air, but burns when heated to whiteness, forming **Oxide of Cadmium**, $CdO=128$.

27. Neither the metal itself, nor its compounds are employed either in medicine or in the arts, except perhaps, the sulphuret, CdS, which has been introduced as a pigment under the name of *cadmium yellow*. Cadmium, like bismuth, is more readily dissolved by nitric, than by any other acid.

EXERCISES ON CHAPTER XXIV.

1. Whence is the name copper derived? Give the symbol and combining equivalent of copper.
2. How is copper found in nature?
3. Whither is the copper ore of Cornwall and Devonshire carried, and why?
4. How is metallic copper obtained from copper pyrites?
5. State the leading properties of copper.
6. How is copper affected by the air?
7. Describe protoxide of copper.
8. Give an account of the suboxide of copper.
9. For what purposes is copper employed?
10. Give the names and composition of the most important alloys of copper.
11. Give the symbol, and physical characteristics of lead.
12. How is lead affected by the atmosphere?
13. How is lead affected by increase of temperature?
14. What is the chief source of metallic lead?
15. How is lead obtained from the ore?
16. What is meant by *dross, slag*, or *scoriæ?*
17. What is massicot, and how is it prepared?
18. What is litharge, and for what is it used?
19. What is red lead?
20. How may peroxide of lead be obtained?

21. How is lead acted on by pure water?
22. When may leaden cisterns be used with safety?
23. Whence is bismuth obtained? and by what process?
24. Name some of the leading properties and uses of bismuth.
25. What is oxide of bismuth, and how may it be prepared?
26. What are the leading properties of cadmium?
27. To what uses has cadmium been applied?

CHAPTER XXV.

METALS WHOSE OXIDES FORM WEAK BASES, OR ACIDS.

TIN, ARSENIC, ANTIMONY, &c.

Tin	*Symbol*	Sn;	*Combining*	*weight*	118.
Arsenic	,,	As;	,,	,,	75.
Antimony	,,	Sb;	,,	,,	122.

1. **Tin** (Latin, *Stannum*) is a comparatively soft and brittle metal, of a silvery-white color, emitting a peculiar odor when rubbed, and producing a characteristic grating or crackling noise when bent; this is probably due to the disturbance of its crystalline structure, and the destruction of cohesion among its particles.

Tin is somewhat more ductile and tenacious than lead, and so malleable, that it may be reduced into leaves (*tin-foil*) $\frac{1}{1000}$th of an inch in thickness.

2. At ordinary temperatures tin is but slightly oxidized, even by the combined action of air and water; but when heated beyond the fusing point (442°) it oxidizes rapidly, being gradually covered with a grey powder, which is the protoxide, SnO. If heated to whiteness in the air, the metal takes fire, and passes into the state of peroxide, SnO_2; if similarly heated in close vessels, it sublimes.

Ex. a.—Let a globule of tin, melted to a white heat, be suffered to fall upon a sheet of paper; it will break into smaller globules, and burst into a bright white flame.

3. Tin, though one of the metals that were earliest known, occurs in but few countries, principally in the mountains of Saxony and Bohemia, the peninsula of Malacca, the island of Banca, and more especially in Cornwall. It was obtained from the latter source by the Phœnicians at a very early period, and the south-west parts of Britain were termed by the ancients the *Cassiterides*, or tin-islands.

4. The principal ore of tin is the peroxide, SnO_2; this is often found associated with copper and zinc ores, in granite or slate rocks, and is known by the name *tinstone*: sometimes it is found in the form of grains, or rounded masses, in alluvial soil, or in the beds of small rivers, in which case it is called *stream-tin*. The former, when reduced to the metallic state, yields *block-* or *bar-tin*; the latter yields *grain-tin*, which is the purer of the two.

5. The ores are reduced to the metallic state by two different methods. In the first and most common process, the roasted ore, mixed with charcoal or culm (*coal-dust*) and a little lime, is carefully heated to redness in a reverberatory furnace so as to fuse the tin, which is let off from time to time into an iron vessel, whence it is ladled into moulds so as to form large blocks.

In the second method, which is only used for the finest ores or *stream-tin*, the ore is fused with charcoal in small blast furnaces, and run off as before.

6. The tin thus obtained has to be refined; this is effected by remelting it in an iron vessel, and continually agitating the mass, either by raising some of it in a ladle, and letting it fall from a considerable height, or by plunging billets of green wood into it. The surface is skimmed from time to time to remove the impurities which collect upon it; the metal is then left at rest awhile, in order that any heavier impurities may settle, after which the upper portion of the tin is ladled into granite, or cast-iron moulds, while the lower portion is remelted with another charge of ore.

tin thus prepared is termed *block-tin*, and till recently was ' to be marked or *coined* with the stamp of the Duchy of l; for a duty is paid to the officers of the Prince of Wales, Duke of Cornwall, on all tin raised in that county.

7. **Peroxide of Tin,** $SnO_2=150$.—This is readily procured by placing metallic tin in slightly diluted nitric acid. Violent action takes place, the acid is decomposed, yielding a portion of its oxygen to the tin, and passing off as nitric oxide, which produces the usual red fumes in contact with the air.

$$3Sn+4HNO_3=3SnO_2+2H_2O+4NO.$$

In general (*as shewn in Chap. VIII.* § 14) a metal when oxidized by the decomposition of nitric acid, unites with another portion of the acid to form a nitrate; not so with tin, for each atom of this metal takes two atoms of oxygen from the acid, and forms a white powder, which is a hydrated peroxide, insoluble in acids. If, however, an alkali be added to a solution of perchloride of tin, a white, hydrated peroxide is precipitated, which is freely soluble in acids. Both hydrates, when moistened, act as acids in reddening litmus paper.

The former of these compounds is called **Metastannic Acid,** the latter, **Stannic Acid,** H_2SnO_3.

Stannate of Sodium, Na_2SnO_3, is much used by dyers and calico printers as a mordant, under the name of *tin-prepare-liquor.*

8. When exposed to a red heat, the hydrated peroxide of tin prepared by either method becomes anhydrous, and is of a straw colour: in this state it is employed by jewellers for polishing, under the name of *putty-powder.* When melted with glass it renders it white and opaque, as in the enamel used for the faces of clocks and watches.

Ex. b.—Keep metallic tin in a state of fusion, and take off the dross or oxide as it forms upon the surface. This dross ground up and washed in water to separate any metallic particles, is the peroxide of tin. Throwing occasionally a little nitre into the crucible will materially facilitate the formation of peroxide.

9. Tin being little affected by air and moisture, is used for coating tacks, pins, wire, and sheets of iron and copper for cooking utensils.

Saucepans, &c., are made of tinned plate, that is, iron coated with tin; this is effected by dipping well-cleaned sheet-iron into melted tin.

Ex. c.—Fill a tinned copper vessel with alternate layers of brass pins, and plates of tin. Cover the whole with a saturated solution of cream of tartar in hot water, and let it boil for five or six hours; when cold, the pins will be found completely coated with tin.

Ex. d.—Clean a slip of copper from all impurities, rub it over with a solution of chloride of ammonium (*sal ammoniac*). Then heat the slip, and immediately rub it over with tallow—now heat it again, and once more rub it over with a piece of tin. This metal will immediately combine with the surface, giving it a silvery coat. The object of the tallow is to prevent oxidation, so that the surface of each metal may continue bright and clean, and not be affected by contact with the air, which would rapidly be the case, at the temperature to which it is raised without this precaution.

Ex. e.—Add a few drops of diluted protochloride of tin to a solution of the perchloride of gold—a purple powder will be thrown down. This powder is used in enamel painting, and for tinging glass of a fine red color; it is a compound of peroxide of tin and oxide of gold, and is termed *purple of Cassius.*

10. Copper and tin form the well-known and highly useful alloys, bell-metal and bronze.

An amalgam of tin and mercury is used for the backs of looking-glasses.

Common pewter is a mixture of tin and lead. Britannia metal, and the best pewter, are tin hardened by a little antimony.

11. **Arsenic** is a brilliant, brittle, steel-grey metal, sometimes found native, but generally in combination with sulphur or oxygen, and associated with iron, nickel, cobalt, lead or silver ores. From these it is separated by heat in a reverberatory furnace. The arsenic combines with oxygen, and is carried off in fumes of **Arsenious acid**, As_2O_3, which are con-

densed in suitable chambers or flues. On heating this arsenious acid with charcoal, it parts with its oxygen to the carbon, and the arsenic is reduced to the metallic form.

12. Arsenic, when raised to about 360°, goes off in vapor without fusing; this vapor has the odor of garlic, and if air be excluded, it readily condenses again in small metallic crystals. It can be fused by heating it in a closed tube under the pressure of its own vapor When heated in the air it easily takes fire, burning with a blue flame, and producing poisonous fumes of white oxide.

13. **Arsenious Anhydride,** $As_2O_3 = 198$.—This substance, which is the common white arsenic of commerce, is a well-known and most deadly poison. It may be obtained either by the actual combustion of the metal, or by sublimation from the various metallic ores which contain it, as already stated.

14. Arsenious anhydride, or *white arsenic*, is a heavy white powder, of a sweetish taste, somewhat soluble in water; the solution is true **Arsenious Acid,** H_3AsO_3. Though sweet it is termed an acid, from its property of neutralizing alkalies and combining with bases, forming compounds termed *arsenites*. Notwithstanding its highly dangerous properties, arsenious acid is much used in dyeing and glass-making, and for agricultural purposes. It prevents the decay of organized substances, hence it is used to preserve timber from *dry-rot*, and to prepare the skins of animals and birds before they are stuffed.

15. The best antidote to arsenious acid is iron-rust (hydrated sesquioxide of iron, Fe_2O_3+3HO), which forms with it a perfectly insoluble compound, that has no poisonous influence. In cases of poisoning by arsenic,—milk, white of eggs, sugar, or soap suds, should be swallowed immediately, and medical assistance procured as speedily as possible.

16. **Arsenic Anhydride,** $As_2O_5 = 230$.—This substance is found in nature associated with various metallic oxides.

It may be formed artificially by oxidizing arsenious acid by means of a heated mixture of nitric and hydrochloric acids, and evaporating the solution to dryness. A white anhydrous powder is obtained, slowly soluble in water, giving a highly acid solution of **Arsenic Acid**, H_3AsO_4.

17. Arsenic very much resembles phosphorus in the characters of its compounds. Both elements unite with oxygen, forming similar acids; these again unite with other bodies, producing compounds perfectly resembling each other in external appearance and crystalline form.*

18. Arsenic also unites with three equivalents of hydrogen to form **Arseniuretted Hydrogen**, AsH_3, a highly poisonous gas analogous to phosphuretted hydrogen, PH_3, already described.

19. There are two compounds of sulphur and arsenic found in nature, and capable of being prepared artificially, viz.—

(1.) **Bisulphuret of Arsenic**, *Realgar*, As_2S, which is obtained by fusing arsenic or arsenious acid with a small proportion of sulphur. It is used in dyeing, and for producing *white fire* in fireworks.

(2.) **Tersulphuret of Arsenic**, *Orpiment*, As_2S_3, made by fusing arsenious acid with excess of sulphur, and sometimes used as a beautiful yellow pigment, termed *King's yellow*.

Ex. f.—Place a very small quantity of white arsenic at the bottom of a dry test tube, and heat it slowly over a lamp. Vapors of arsenious acid will arise and condense in brilliant *octohedral* crystals, the regular triangular faces of which may be observed with a magnifying glass.

Ex. g.—Take a glass tube having a bulb at the end, or draw out a test tube in the flame of a lamp so that the lower part may serve as a bulb. Place in the bulb a few particles of white arsenic mixed with some cyanide of potassium, or with a little charcoal and dried carbonate of sodium; on heating the mixture, the arsenious acid will part with its

* Bodies thus closely resembling each other are termed *isomorphous*.

oxygen to the cyanide, or the charcoal, and metallic arsenic will condense in a brilliant ring on the narrow part of the tube above the bulb.

Ex. h.—The adjoining diagram represents Dr. Marsh's apparatus for testing the presence of arsenic. A few pieces of *pure* zinc are placed in the bulb below the stop-cock, and the solution supposed to contain arsenic (previously mixed with *pure* sulphuric acid) is poured in at the mouth of the apparatus. Hydrogen gas is formed (Chap. v. § 11), which combines with the arsenic, forming arseniuretted hydrogen, AsH_3.

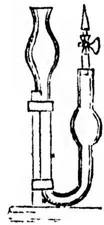


Fig. 71.

$$As_2O_3 + 12H = 3H_2O + 2AsH_3.$$

The gas accumulating in the shorter limb, forces the liquid past the zinc into the longer tube, and the action ceases. When the stopcock is opened, the gas is forced out by the pressure, and if ignited burns with a white flame, producing fumes of arsenious acid; but if a cold body, such as a piece of glass, earthenware, or porcelain, be held over the jet of burning gas, thin films of metallic arsenic will be deposited in the form of brilliant rings.

Ex. i.—Dissolve some arsenious acid, by boiling it with a few drops of hydrochloric acid, and pass through the solution a little sulphuretted hydrogen. The sulphur of the gas combines with the metal of the arsenious acid, forming a bright yellow precipitate of orpiment (*King's yellow*), As_2S_3, while the hydrogen combines with the liberated oxygen to form water.

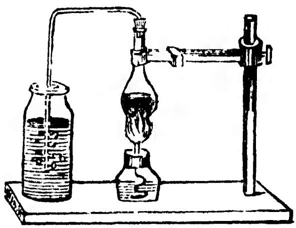


Fig. 72.

$$2(H_3AsO_3) + 3H_2S = As_2S_3 + 6H_2O.$$

The yellow precipitate is soluble in ammonia, and may be reprecipitated by acids.

Ex. k.—Drop a solution of nitrate of silver into a solution of arsenious acid in distilled water. No precipitate is thrown down, nitric acid having a stronger affinity for oxide of silver than arsenious acid has. If a little potass or ammonia be now added, it combines with the nitric acid, and forms nitrate of potassium, or ammonium, which remains in solution; while the arsenious acid combining with the oxide of silver, forms a yellow-coloured precipitate—the arsenite of silver, Ag_3AsO_3, thus:—

$$6AgNO_3 + 2H_3AsO_3 + 3K_2O = 6KNO_3 + 2Ag_3AsO_3 \quad H_2O.$$

Ex. l.—Use sulphate of copper instead of nitrate of silver in the preceding experiment, and the result will be a green precipitate of arsenite of copper, (*Scheele's green,*) $CuHAsO_3$, thus :—

$$CuSO_4 + H_3AsO_3 + 2NH_3 = CuHAsO_3 + (NH_4)_2SO_4$$

Ex. m.—Mix together six parts of saltpetre, two of sulphur, and one of realgar or orpiment. Touch the mixture with a red-hot wire, it will burn with a vivid white light (*Bengal light*).

Ex. n.—Throw a few grains of powdered metallic arsenic into a red-hot crucible, it will burst into a bright flame.

All experiments in which vapors of arsenic arise should be performed out of doors, or on the hob by the side of the fire, so that the draught of air up the chimney may carry them away.

20. **Antimony** is a brilliant metal, of a bluish-white color, extremely brittle, and possessing a crystalline structure. It is nearly seven times heavier than water (6.7).

21. Antimony fuses at 800°, and may be slowly volatilized at a white heat, out of contact with the air.

It remains unacted upon by the air at ordinary temperatures, but is rapidly oxidized when exposed to it in a state of fusion.

Ex. o.—Perform *Ex. e.* Chap. xxiii. with antimony instead of zinc, the combustion will be very vivid, and a white smoke of oxide of antimony, Sb_2O_3, will be given off.

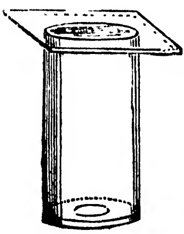


Fig. 73.

Ex. p.—Reduce some antimony to powder, and heat it to about 80°, then pour it into a jar of chlorine. The metal will ignite spontaneously, burning with a white flame, and giving rise to chloride of antimony, $SbCl_3$.

22. Native antimony is found in France, Germany, and Sweden, but the metal is chiefly obtained from the sulphuret, Sb_2S_3, which occurs in the mountains of Auvergne, in the Hartz mountains, at Schemnitz in Hungary, and in the south-west of Cornwall.

23. Metallic antimony may be obtained from the sulphuret, by simple fusion with iron, in earthen crucibles.

The iron, having a greater affinity for sulphur than antimony has, unites with it, forming sulphuret of iron, which floats as a slag, while the heavier antimony settles to the bottom of the crucible.

24. **Oxide of Antimony, Sb_2O_3** = 292, has been already referred to as the product formed by fusing metallic antimony in a current of air. An impure oxide, formerly called *glass of antimony*, is obtained by roasting the sulphuret in a reverberatory furnace, till all the sulphur is expelled.

25. When oxide of antimony is boiled with a solution of tartrate of potass (*cream of tartar*), a compound of tartaric acid,* potass, and antimony is formed, which readily crystallizes on evaporation.

This compound is known in medicine as *tartar emetic*, and is a tartrate of potass and antimony, $K,SbO,C_4H_4O_6$.

26. If oxide of antimony be heated in a crucible, open to the air, it combines with more oxygen, forming a greyish-white, infusible powder, insoluble in water or acids, but capable of uniting with alkalies and forming solutions, from which acids precipitate **Antimonious Acid**, $H_2Sb_2O_5$. The white powder is **Antimonious Anhydride**, Sb_2O_4 = 308.

27. When metallic antimony is added to hot nitric acid, the acid is decomposed, as in the case of tin, § 7, and an insoluble, pale yellow powder is formed, which is the highest oxide of antimony, and is termed **Antimonic Anhydride**, Sb_2C_5 = 324.

28. **Sulphuret of Antimony**, Sb_2S_3 = 340.—This compound, which is often found native, is a brittle, crystalline substance of a leaden-grey color, and very fusible. When artificially prepared, (as it may be, by adding sulphuretted hydrogen or sulphuret of ammonium to a solution of tartar-emetic or other soluble compound of antimony,) it is precipitated of an *orange-yellow* colour, but darkens on drying,

* The formula for tartaric acid or T', is $H_2C_4H_4O_6$.

It is also found somewhat abundantly in Bohemia, Hungary, Sweden, China, Japan, Mexico, and Chili.

3. Mercury has a silvery-white color, and a brilliant lustre; it is distinguished from all other metals by being liquid at ordinary temperatures. When cooled to —39° F., (or 71 degrees below the freezing point of water,) it solidifies, and is then soft, malleable, and ductile. At a still lower temperature, however, it becomes brittle. It is about 13½ times as heavy as water, (sp. grav. 13.56).

4. Mercury emits vapors at all temperatures above its freezing point (—39° F.), but they are scarcely perceptible till the heat rises to 68° or 70°. At 660° mercury boils, giving off a transparent, colorless vapor, which readily condenses on cooling.

Ex. a.—Pound some vermilion, HgS, mix with it an equal weight of charcoal powder; place this mixed powder in an earthenware retort, and submit it to a red heat. The metal will sublime, and run down the neck of the retort into any vessel placed to receive it.

5. Pure mercury is quite unaltered in the air, at common temperatures, but when heated to near its boiling point, and exposed to the action of air, it absorbs oxygen, and becomes changed into a dark red, crystalline powder, which is the *red* or *per*-oxide of mercury.

"Nitric acid dissolves mercury, yet not so readily as it does the metals tin, lead, and zinc, with which mercury is often contaminated, either naturally, accidentally, or for the purpose of adulteration. To remove these, the impure mercury may be placed in a phial along with a *small* quantity of nitric acid, and shaken up; this acid will immediately attack the solid metals. When they are dissolved, the solution may be poured off, and the mercury washed in water to remove any adherent acid, it will then be quite pure."—*Francis.*

6. Mercury forms two compounds with oxygen, viz. a suboxide (or *protoxide*) and an oxide (or *peroxide*).

The combining weight of mercury is given differently by different chemists, some thinking it 100, others 200, hence the precise nomenclature and combining weights of its compounds are somewhat unsettled.

7. **Suboxide of Mercury,** or *grey oxide,* $Hg_2O=416$.—This is a dark-grey or nearly black powder, very heavy, and insoluble in water, but uniting with acids to form salts. It may be obtained by adding potass, soda, or lime-water to a *sub-salt* of mercury, such as the sub-nitrate (or *proto-nitrate*) $HgNO_3$, or the subchloride (*calomel*), HgCl. Thus—

$$2HgNO_3 + \left\{\begin{matrix} K_2O \\ Na_2O \\ CaO \end{matrix}\right\} = \left\{\begin{matrix} 2KNO_3 \\ 2NaNO_3 \\ Ca2NO_3 \end{matrix}\right\} + Hg_2O.$$

8. **Oxide of Mercury,** or *red oxide,* $HgO=216$. This substance has been already mentioned as a dark-red, crystalline powder, produced by the action of the air on heated mercury. It has a yellow tint when finely pulverized, and if the temperature be slightly raised, it turns black, becoming red again on cooling. The red oxide of mercury may also be obtained by carefully heating the nitrate, $Hg2NO_3$, so as to expel the nitric acid; or by adding potass, soda, or lime-water to a solution of the chloride (*corrosive sublimate*), $HgCl_2$. In the latter case the precipitate is of a yellow color, being in a state of minute division. If strongly heated, the red oxide of mercury is decomposed into oxygen and metallic mercury, which will go over in vapor, as shewn in Chap. IV. § 4.

In this way the presence of impurities, as red-lead, brick-dust, &c., may be detected, for they will remain in the retort or flask.

9. **Sulphuret of Mercury,** $HgS=232$.—This substance is found native; it is the most common ore of mercury, and known as *cinnabar,* or vermilion. It may be artificially prepared by subliming a mixture of six parts of mercury and one of sulphur, and afterwards reducing the mass obtained to a fine powder. Vermilion is inodorous, tasteless, and insoluble. It is unaltered by exposure to the air or moisture, but when heated to dull redness, it is resolved into vapor of mercury and sulphurous acid gas, SO_2.

Ex. b.—Rub together in a mortar three parts of sulphur and one of mercury. The lustre of the one, and the yellow color of the other,

will disappear, the whole being converted into a black powder, which was formerly called *Ethiops* mineral; and is probably a mixture of the sulphuret of mercury with additional sulphur.

10. Mercury is extensively used in the manufacture of barometers and thermometers, also for "silvering" mirrors. In the latter case an *amalgam* of tin is employed; an amalgam of silver is sometimes used by dentists for filling up hollow teeth.

Ex. c.—Melt two ounces of bismuth and two ounces of lead in separate crucibles, pour them into separate vessels, containing an ounce of mercury in each; when cold, these alloys will be in a solid state; but if they are rubbed against each other, they will instantly become liquid.

Ex. d.—Place a little gold-leaf in the palm of the hand, and pour upon it a globule of mercury. The latter will be seen to absorb, or combine with the gold, forming a more or less fluid and yellow amalgam, according to the proportion of the two metals. This amalgam is used in *water gilding*. The affinity of mercury for gold and silver is so strong, that persons who are foolish enough to clean their watch cases with mercury, or one of its salts, will find them totally spoiled; the same holds good with respect to plated articles cleaned by a composition, sold about the streets, made of the nitrate of mercury, ground up with whitening. Even those who are obliged to take much calomel, or other medicines containing mercury, should abstain from wearing any gold articles, or carrying gold money, as the mercury sometimes oozes through the pores of the skin, and attaches itself to the gold, rendering it so brittle, that it may often, when thus contaminated, be broken in two. The best way of restoring money thus spoiled is to heat it red-hot in a crucible, or ladle.

Ex. e.—Add five parts of mercury to two parts of melted zinc and one of tin, an amalgam will be formed such as is used for the cushions or rubbers of electrical machines. The tin may be omitted.

11. **Silver,** (Latin, *Argentum*, whence the symbol Ag,) is the whitest of the metals; it has a very brilliant lustre, and will take a very high polish. It is, when pure, soft enough to be cut with a knife, being thus softer than copper, and harder than gold. It is about 10½ times as heavy as water; its specific gravity being 10.4 when cast, and 10.5 to 10.6 when stamped or rolled.

12. Silver is exceedingly malleable and ductile, for it may be extended into leaves less than $\frac{1}{100000}$th of an inch in thickness, and a single grain may be drawn into a wire 400 feet long. Only three metals, viz. iron, copper, and platinum, surpass it in tenacity, a wire $\frac{1}{10}$th of an inch in diameter being able to support 187lbs. without breaking.

13. Silver does not rust or oxidize by exposure to the air or moisture; but when the atmosphere contains sulphureous vapors, or sulphuretted hydrogen, H_2S, the metal speedily tarnishes, owing to its affinity for sulphur, becoming first yellow and finally black.

Ex. e.—See *Ex. o* Chap. XIV.

14. Silver fuses at about 1873° Fahr., and during the state of fusion, absorbs considerable quantities of oxygen from the air, *which it parts with again on cooling.*

When very strongly heated in a blast furnace, silver rises in vapor; and when a strong electric current is made to pass through some leaf-silver or a fine silver wire, the metal burns with a greenish flame.

15. Silver, it has been already stated, does not oxidize when heated alone, but if heated with silicious matter, such as glass, porcelain, &c., it combines with oxygen, and forms an oxide of silver, $2Ag_2O$, that immediately unites with the silicic acid or silica, SiO_2, forming a yellow silicate of silver, Ag_4SiO_4.

16. **Oxide of Silver**, $Ag_2O = 232$.—This oxide is employed in glass, porcelain, or enamel painting, to produce a rich yellow tint; it is also the base of all the common *salts* of silver. One of the most characteristic features of these compounds, is their property of darkening by exposure to light, whence their use in all branches of Photography.

If caustic potass, KHO, be added to a solution of ni-

trate of silver, $AgNO_3$, a hydrated oxide of silver will be precipitated.

$$AgNO_3 + KHO = KNO_3 + AgHO.$$

17. Silver is sometimes found in the native state, sometimes alloyed with other metals, but chiefly in combination with sulphur or chlorine. The principal silver mines are to be found at Konigsberg in Norway, among the Hartz mountains in Germany, and among the Andes of Mexico and Peru.

18. Metallic silver may be obtained from the sulphuret or chloride by the process of *amalgamation.* The powdered sulphuret is mixed with common salt (chloride of sodium), and roasted at a low red-heat in order to convert it into the chloride. The mixture thus formed, or the native chloride, (if that ore be employed,) is then crushed and agitated for some time with mercury, water, and scraps of iron, by which means the chloride of silver is decomposed; the protochloride of iron formed is dissolved by the water, while the mercury combines with the reduced silver, (as well as with any gold, copper, or lead, that may be present,) to form an amalgam. This amalgam is then pressed in woollen bags to remove any uncombined mercury, and heated in an iron retort, whereby the mercury is distilled off, and an impure silver left behind, which afterwards undergoes the process of refining.

19. Almost all lead ores, but especially *galena,* a sulphuret of lead, contain a proportion of silver; and it has been found profitable to extract this, even when so low as eight ounces in the ton.

The *argentiferous* or silver-bearing lead, reduced from the ore as shewn in a former chapter, is melted and then allowed to cool slowly; the lead begins to solidify first, and while doing so, is removed, from time to time, with a perforated iron ladle or strainer. An alloy containing a large proportion of silver remains in the liquid state; this is allowed to cool, and the lead is afterwards separated by *cupellation.*

20. The process of *cupellation* consists in exposing the alloy of silver and lead on a *cupel* in a furnace to a current of heated air, whereby the latter metal is oxidized. The *cupel* is a kind of shallow, porous, crucible made of bone-ash, the object of which is not only to support the metal, but to absorb the oxide of lead as fast as it is formed.

21. Pure silver is too soft to be used with advantage, it is therefore generally alloyed with copper. The standard silver used for the English coinage consists of thirty-seven parts pure silver, and three parts copper.

Ex. f.—Dissolve a silver coin in nitric acid, the solution will be of a greenish color, owing to the presence of nitrate of copper; place a piece of pure copper in the solution the acid will leave the silver and dissolve the copper, the former metal being precipitated as a crystalline powder, which, when well washed and dried, may be fused into a button of pure silver.

Ex. g.—Instead of immersing a piece of copper in the above solution of nitrate of silver, add a little salt and water so long as a precipitate is formed. This precipitate is chloride of silver, AgCl, which, if washed, dried, and melted with carbonate of sodium, will yield chloride of sodium and pure silver.

$$2AgCl + Na_2CO_3 = 2NaCl + CO_2 + O + 2Ag.$$

22. **Gold** (*Aurum*) is one of the most widely diffused of the metals, being found in almost every country, but most abundantly in South America, California, Australia, the East Indies, Western Africa, Hungary, and the Uralian mountains.

23. Gold is found exclusively in the metallic state, seldom indeed pure, but alloyed either with silver, copper, or some other metal. It is occasionally met with in masses or nuggets of several ounces or even pounds in weight; but it occurs more generally in the form of grains, disseminated through quartz and other rocks, or in the sands of various rivers.

24. These gold-sands (the *gold-dust* of commerce) are derived from the *disintegration* or breaking up of the rocks in which the metal was contained. The gold, owing to its

high specific gravity (19.2), is readily separated from the sandy or earthy matter by washing in water; or it may be obtained by the process of *amalgamation*, as in the case of silver.

25. Gold is the most malleable and ductile of the metals, and likewise very tenacious.

It may be extended into leaves less than $\frac{1}{280000}$th of an inch in thickness;—a single grain may be drawn into a wire 500 feet in length;—and a wire $\frac{1}{10}$th of an inch in diameter can support a weight of 158 lbs. (See Chapter XIX., § 10).

26. Gold does not tarnish or oxidize when exposed either to heat, air, or moisture. It is unacted on by any single acid, but is dissolved by chlorine, especially in the *nascent* state, for which purpose a mixture of nitric and hydrochloric acids is used, termed nitro-muriatic, or nitro-hydrochloric acid, or *aqua regia*. The resulting compound is **Perchloride of Gold**, $AuCl_3$.

Ex. h.—Place some gold-leaf in a glass of water, and pass a current of chlorine through the liquid, the metal will be rapidly dissolved, forming perchloride of gold. The same compound may be produced by dissolving leaf-gold in nitro-muriatic acid.

27. Pure gold, being very soft, as well as very costly, is never employed in the arts. It is generally alloyed with silver or copper, to give it greater hardness. The metal used for English gold coin consists of eleven parts gold, and one part alloy.

28. **Platinum** derives its name from the Spanish *plata*, silver, on account of its silvery lustre. It is about half as valuable as gold, and is always found in the metallic state, occurring in the form of flattened grains of a greyish-white color, and associated with palladium, osmium, iridium, rhodium, and a little iron.

29. Platinum is chiefly obtained from alluvial deposits

in the Uralian mountains :* it is also found in St. Domingo, Borneo, and Ceylon, and in the gold districts of South America. It is separated from the sandy and earthy matter with which it is associated, by washing in water, as in the case of gold and stream-tin.

30. Pure platinum is very ductile and malleable, and one of the most tenacious of the metals. It is the heaviest substance known, being nearly 22 times as heavy as water. Its specific gravity varies from 19.5 to 21.9.

31. Platinum is unchanged by exposure to air or moisture. It remains infusible in the strongest heat of a smith's forge, and can only be melted by the agency of electricity, or by the oxy-hydrogen blowpipe. Like iron, it possesses the property of *welding*, when heated to whiteness. (Chap. XXIII., § 14.

32. Platinum is insoluble in any of the simple acids; but, like gold, it may be dissolved in *aqua regia*, or nitro-hydrochloric acid, giving rise to **Bichloride of Platinum**, $PtCl_4$.

Platinum is speedily oxidized and corroded by the caustic alkalies at a red heat; it likewise forms alloys with most of the other metals.

Ex. i.—Place a particle of lead, tin, or antimony on a piece of platinum foil, and heat it over a spirit lamp; when at a red heat, the metal used, whichever it may be, will combine with the platinum, and form a fusible alloy. This experiment shows that the above metals must never be melted in platinum vessels; the same holds good with the ores and oxides of these metals, and in a less degree with bismuth, copper, cobalt, and nickel. The following cautions are added by Berzelius :—"Platinum vessels must not be subjected at any time to the action of chlorine, or the compounds which evolve chlorine. Nitre and the alkalies must not be fused in them. Immediate contact with the fuel, which ought to be charcoal, (if in a furnace,) must be avoided as much as possible."

* The largest specimen yet discovered was found in the Uralian mountains, and weighs 21 lbs. troy.

33. **Platinum may be obtained in two states of very minute division; in the one state it is termed *spongy platinum*; in the other, which is still finer, it is called *platinum black*. Each of these substances has the property of condensing gases in its pores. The result of this condensation is to bring the gaseous bodies in close contact, and cause them to combine. In this way oxygen and hydrogen are compelled to unite, and flame is produced, as in *Dobereiner's* lamp.**

Ex. k.—Spongy platinum may be made by saturating blotting paper with bichloride of platinum and drying it, repeating the process several times, and finally, burning the paper in the flame of a spirit lamp; the ashes left are the spongy platinum.

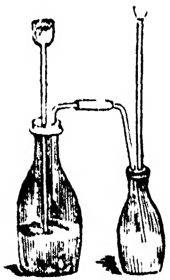


Fig. 74.

Ex. l—To a diluted solution of bichloride of platinum add some solution of chloride of ammonium (sal-ammoniac), a yellow precipitate of ammonio-chloride of platinum, $PtCl_2+NH_4Cl$, is produced, which is decomposed by heat, and yields a grey mass of spongy platinum; this becomes red hot when exposed to a stream of hydrogen, and then inflames the gas. (*Doberei-*

Ex. m.—The peculiar property of this metal in continuing red hot, when exposed to certain vapors, is shown not merely by spongy platinum, but by platinum in its ordinary metallic state. This may be tried as follows:—Make a small coil of *fine* platinum wire, and place it loosely round the wick of a spirit lamp, filled with alcohol; set fire to the wick, and when it has burnt two or three minutes, blow out the flame, the platinum wire will continue red hot until the whole of the spirit be exhausted. It should be covered loosely with a glass funnel, so that air may have access to it, yet no current be sufficient to rob the platinum of its heat.

34. **Platinum being infusible, cannot be obtained from its ore, by furnace operations, in the usual way; it is therefore dissolved in *aqua regia*, and the resulting chloride precipitated by sal-ammoniac, as in Ex. *l.*; this, when heated, yields the metal in a finely divided state. The spongy platinum formed is then heated to whiteness,**

and its particles welded together by strong pressure. On being again heated, it can be hammered or rolled into plates, drawn into wire, &c.

35. Platinum, being unacted on either by acids or by heat, is employed in the manufacture of crucibles and other chemical apparatus. Large stills made of this metal are used in the concentration of sulphuric acid. (Chap. XIV. § 21.) It is also used by gun-makers for making the touch-holes or nipples of fire-arms; and in Russia, it has been coined into money.

EXERCISES ON CHAPTER XXVI.

1. In what form is mercury met with in nature?
2. Whence is mercury obtained?
3. Give the symbol, combining equivalent, specific gravity, and leading properties of mercury.
4. How is mercury affected by change of temperature?
5. How is mercury affected by the air?
6. What compounds does mercury form with oxygen, and why is their nomenclature unsettled?
7. Describe the grey oxide of mercury, and the modes of obtaining it.
8. How may the red oxide of mercury be obtained, and what are its properties?
9. State the nature and properties of vermilion.
10. For what is mercury used?
11. What are the physical characteristics of silver?
12. Illustrate the malleability, ductility, and tenacity of silver.
13. How is silver affected by the atmosphere?
14. How is silver affected by heat?
15. How is silicate of silver formed, and to what purpose is it applied?
16. For what are oxide of silver and its compounds used?
17. Whence is silver obtained?
18. How may metallic silver be obtained from the sulphuret or chloride?
19. How may silver be obtained from the lead ores in which it is frequently met with?
20. What is the process of *cupellation?*
21. What is standard silver?
22. Whence is gold obtained?
23. Under what forms is gold found?

24. What is the origin of gold-sands, and how may the metal be obtained from them?
25. Illustrate the ductility, tenacity, and malleability of gold.
26. How is gold affected by heat, air, moisture, and acids? How may gold be dissolved?
27. Why is gold alloyed? What is standard gold?
28. Whence does platinum take its name, and how is it found in nature?
29. Whence is platinum obtained?
30. State some of the physical properties of platinum.
31. How is platinum affected by heat?
32. How is platinum dissolved?
33. State the nature and properties of spongy platinum and platinum-black.
34. How is platinum obtained from its ores?
35. For what is platinum employed?

CHAPTER XXVII.

SALTS.

1. MOST of the compounds of the metalloids with oxygen are **acidifiable oxides**, which, when combined with water, form the class of acids formerly called *oxygen-acids* (Chap. XIX. § 14).

2. Most of the compounds of the metalloids with hydrogen are **acids**, and were formerly termed *hydracids* (Chap. VII. § 8).

3. Most of the compounds of the metals with oxygen are **bases**, or *basic oxides* (Chap. XIX. § 15).

4. When acids and bases combine, they more or less completely neutralize each other, and form a class of compounds termed **Salts**.

5. Salts may also be defined to be compound bodies produced, either.

(i.) By the union of an acid with a base, as nitrate of silver, sulphate of copper, carbonate of calcium, &c.

Or, (ii.) By the union of a metal with a metalloid, as chloride of zinc, iodide of potassium, &c.

6. Salts of the latter class have been termed **Haloid Salts**, from their analogy or resemblance to sea-salt (ἅλς, *hals*, salt; and εἶδος, *eidos*, form).

7. Salts formed from acids whose names end in *-ic*, are known by the ending *-ate*; those formed from acids ending in *-ous* have names ending in *-ite*.

Thus, nitr*ic* acid forms nitr*ates*; nitr*ous* acid, nitr*ites*; sulphur*ic* acid, sulph*ates*, &c.

8. At the present time all acids are considered to be compounds of hydrogen; in fact they are hydrogen-salts in which the hydrogen acts the part of a metallic base (Chap. VII. § 12).

9. According to this theory, a salt is formed by the replacement of the whole or part of the hydrogen in the acid by its equivalent in metal.

10. When a mono-basic acid (Chap. XVIII. § 20) acts on a mono-atomic metal, the single atom of replaceable hydrogen is replaced by an atom of the metal.

Thus, chloride of sodium (NaCl) may be obtained by acting on hydrochloric acid (HCl) by sodium (Na); the atom of hydrogen in the acid being replaced by an atom of the metal.

$$HCl + Na = NaCl + H.$$

Again, chloride of sodium may be obtained by adding one atom of soda (Na_2O) to two atoms of hydrochloric acid (2HCl); the two atoms of sodium replacing the two atoms of hydrogen, which combine with the oxygen to form an atom of water.

$$2HCl + Na_2O = 2NaCl + H_2O.$$

When a mono-basic acid acts on a di-atomic metal, two atoms of acid must be taken for each atom of metal.

For example, chloride of barium (a di-atomic metal) may be formed by the action of hydrochloric acid on metallic barium, on caustic

baryta or oxide of barium, or on the hydrated oxide of barium; thus:—

$$2HCl + Ba = BaCl_2 + H_2.$$
$$2HCl + BaO = BaCl_2 + H_2O.$$
$$2HCl + BaH_2O_2 = BaCl_2 + 2H_2O.$$

11. Conversely, when a di-basic acid acts on a mono-atomic metal, two atoms of metal must be taken to form a *neutral salt.*

For example, sulphate of potassium (a mono-atomic metal) is formed by the action of sulphuric acid (a di-basic acid) on potassium, on caustic potass, or on the hydrate of potassium; thus:—

$$H_2SO_4 + 2K = K_2SO_4 + H_2.$$
$$H_2SO_4 + K_2O = K_2SO_4 + H_2O.$$
$$H_2SO_4 + 2KHO = K_2SO_4 + 2H_2O.$$

12. A di-basic acid is neutralized by a di-atomic metal atom for atom. Thus,

$$H_2SO_4 + Ca = CaSO_4 + H_2.$$
$$H_2SO_4 + CaO = CaSO_4 + H_2O.$$
$$H_2SO_4 + CaH_2O_2 = CaSO_4 2H_2O.$$

13. In order to form a *neutral salt* with a tri-basic acid, we must take, for a

Mono-atomic metal,	3 atoms of	metal to	1 atom of	acid.
Di-atomic	,,	3 ,,	,,	2 atoms	,,
Tri-atomic	,,	1(3) ,,	,,	1(3) ,,	,,

Thus, Phosphate of Sodium Na_3PO_4.
Phosphate of Calcium Ca_32PO_4.
Phosphate of Bismuth = $Bi_33PO_4 = BiPO_4$.

14. Conversely, to form a *neutral salt* with a tri-atomic metal, we must have,

1 atom of	metal to	1 atom of	a	tri-basic acid.
2 atoms	,,	3 atoms	,,	di-basic ,,
1 atom	,,	3 ,,	,,	mono-basic acid.

Thus, Phosphate of Bismuth = $BiPO_4$.
Sulphate of Aluminum = Al_23SO_4.
Nitrate of Bismuth = $Bi3NO_3$.

15. **When only part of the hydrogen in the acid is replaced by a metal, the resulting compound is an *acid salt* or *super-salt*.**

For instance, in sulphuric acid (H_2SO_4), which is di-basic, one or both of the atoms of hydrogen may be replaced by potassium. Thus,

Sulphuric Acid	... ...	= H_2SO.
Neutral Sulphate of Potassium	... ...	= K_2SO_4.
Acid Sulphate, or Per-sulphate of Potassium		= $KHSO_4$.

Or, again:—In phosphoric acid (H_3PO_4), which is tri-basic, one, two, or three atoms of hydrogen may be replaced by sodium. Thus,

Phosphoric Acid	... ...	= H_3PO_4.
Basic Phosphate of Sodium	..	= Na_3PO_4.
Neutral Phosphate of Sodium . .		= Na_2HPO_4.
Acid-, or Per-phosphate of Sodium		NaH_2PO_4.

16. **The metallic or quasi-metallic half of a salt is called the *base* or *basyle*.**

The non-metallic part of a salt is sometimes termed the *salt radical*.

Salt.	Symbol.		Base.	Salt-radical.
Thus, Chloride of Sodium	$NaCl$	=	Na	Cl.
Sulphate of Iron	$FeSO_4$	=	Fe	SO_4.
Carbonate of Calcium	$CaCO_3$		Ca	CO_3.
Nitrate of Ammonium	$(NH_4)NO_3$	=	(NH_4)	NO_3.
Cyanide of Potassium	$K(CN)$	=	K	(CN).

17. ***Acid* salts, or *per* salts, in which the hydrogen of the acid has not been entirely replaced by the base, or in which the salt-radical is in excess of the base, are frequently marked by the Latin prefixes, *bi-*, *ter-*, *sesqui-*, &c.; as—**

Bi-sulphate of Potassium	$KHSO_4$.
Bi-carbonate of Sodium	$NaHCO_3$.
Bi-sulphuret of Iron	FeS_2.
Ter-borate of Potassium	KH_23BO_3.

Or, again, Chromate of Potassium $\left.\begin{matrix} K_2 \\ CrO_3 \end{matrix}\right\} O = K_2CrO_4.$

Bichromate " $\left.\begin{matrix} K_2 \\ 2CrO_3 \end{matrix}\right\} O = K_2Cr_2O_7.$

Terchromate " $\left.\begin{matrix} K_2 \\ 3CrO_3 \end{matrix}\right\} O = K_2Cr_3O_{10}.$

18. The prefixes *proto-*, *per-*, &c., before the name of an oxygen-salt, show that the base is in the state of *proto-*oxide, *per-*oxide, &c.

Thus, proto-sulphate of iron, $FeSO_4$, is the sulphate of the protoxide (FeO,SO_3).

Per-sulphate of iron, $Fe_2 3SO_4$, is the sulphate of the peroxide ($Fe_2O_3,3SO_3$.

19. *Basic*, or *sub* salts, in which the base is in excess of the salt-radical are frequently marked by the Latin prefix *sub-*, or the Greek prefixes *di-*, *tri-*, &c.

Thus, Sub-chloride of Mercury $HgCl$.
Sub-oxide of Copper Cu_2O.
Sub-carbonate of Copper $CuCO_3,CuH_2O_2$.

20. The replaceable atoms of hydrogen in an acid may be replaced by different metals according to their equivalent power, hence we have *double salts*.

For instance, *common alum* is a sulphate of aluminum and potassium, four atoms of hydrogen in two atoms of acid being replaced by one of tri-atomic aluminum, and one of mono-atomic potassium. Thus,

$$\left.\begin{matrix} H_3 \\ H \end{matrix}\right\} 2SO_4 \text{ becomes } \left.\begin{matrix} Al \\ K \end{matrix}\right\} 2SO_4 = AlK2SO_4.$$

Again, in the double sulphate of copper and potassium,

$$\left.\begin{matrix} H_2 \\ H_2 \end{matrix}\right\} 2SO_4 \text{ becomes } \left.\begin{matrix} Cu \\ K_2 \end{matrix}\right\} 2SO_4 = CuK_2 2SO_4.$$

Again, in the (tri-basic) phosphoric acid series, we have a double phosphate of magnesium and ammonium, for

$$\left.\begin{matrix} H_2 \\ H \end{matrix}\right\} PO_4 \text{ becomes } \left.\begin{matrix} Mg \\ (NH_4) \end{matrix}\right\} PO_4 = Mg\,NH_4\,PO_4.$$

Once more, in the mineral called *dolomite*, we have a double carbonate of calcium and magnesium (both di-atomic metals). Thus,

$$\left.\begin{matrix} H_2 \\ H_2 \end{matrix}\right\} 2CO_3 \text{ becomes } \left.\begin{matrix} Ca \\ Mg \end{matrix}\right\} 2CO_3 = CaMg2CO_3.$$

Ex. a.—Pour a strong solution of oxalic acid, $H_2C_2O_4$, on some copper filings, an oxalate of copper, $Cu(C_2O_4)$, is formed; if ammonia be added, decomposition ensues, oxalate of ammonia is produced, and oxide of copper precipitated.

$$Cu(C_2O_4) + 2(NH_4)HO = (NH_4)_2(C_2O)_4 + H_2O + CuO.$$

Ex. b.—If instead of pure ammonia we add oxalate of ammonia to oxalate of copper, no decomposition takes place, but a double salt called oxalate of ammonia and copper is the result, $Cu(NH_4)_2 + 2(C_2O_4)$. This double salt forms *blue rhomboidal* crystals, while the separate salts appear,—the one, the oxalate of ammonia, in the state of *flat, white, prismatic* crystals; the other, the oxalate of copper, in the form of a *green* powder.

21. Many of the *acid* or *super-* salts may be strictly considered double salts, in which *water forms one of the bases*:—thus common bisulphate of potassium, $KHSO_4$, might be regarded as a double sulphate of potassium, and water, $K_2H_2 2SO_4 = 2KHSO_4$.

22. The particles of almost every substance, simple or compound, have, when passing from the fluid to the solid state, a tendency to arrange themselves in certain geometrical figures, usually bounded by flat surfaces. These figures are termed *crystals*, and the process by which they are formed is termed *crystallization*.

Thus common salt crystallizes in the form of cubes or dice; Epsom salts, in four-sided prisms; salt petre, in six-sided prisms; alum, in octohedra or double pyramids; &c.

23. Crystallization takes place during the passage of bodies from the fluid to the solid state. This may be effected in several ways.

First, by simple fusion and slow cooling, as in the case of sulphur, the metals, &c.

See Chap. XIV., *Ex. a*; Chap. XXIV., *Ex. h*.

Secondly, by sublimation, as in the case of iodine, arsenic, and antimony, camphor, naphthaline, benzoic acid, &c.

See Chap. XVI., *Ex. m*; Chap. XXV, *Exs h.* and *q*.

Thirdly, by solution and subsequent evaporation, as in the case of most salts and organic compounds.

Ex. c.—Make a strong solution of loaf sugar, pour the syrup thus made into a vessel of any kind, which place in a slow oven; the

evaporation being rapid, the sugar will be deposited in very small crystals, or, in fact, the whole will become loaf sugar again. Place some of the same solution in a pan, stretch one or more strings across the vessel and set it aside, so that the evaporation may be very slow; instead of the sugar presenting a fine granular surface as before, it will appear in large well-formed crystals clinging to the strings. In this state it is known as sugar candy.

24. It is a general rule that every crystallizable substance has its own crystalline form. The same body may, however, under different circumstances be found crystallised on two different systems,* it is then termed *dimorphous* (two-formed); thus carbon appears in octohedral crystals in the diamond, and in hexagonal or six-sided plates in plumbago; carbonate of calcium and sulphur are other examples of dimorphous bodies.

25. When different substances of similar chemical constitutions crystallize in similar forms, and can consequently replace each other without altering the geometrical figure of the crystal, they are termed *isomorphous*, (equal or like-formed). Crystals of isomorphous bodies, when of the same colour, cannot be distinguished by the eye; this is the case, for instance, with the sulphates of zinc and magnesium.

26. Most crystallized salts contain what is termed *water of crystallization*. This water, which forms *part of the crystal, not of the salt*, may be removed by heat or by exposure to a dry atmosphere, and must be distinguished from the basic water essential to the existence of certain salts as noticed, § 21.

27. Some crystallized salts are naturally *anhydrous*, that is, they contain no water of crystallization, as nitrate of potassium (*saltpetre*).

Others become anhydrous when heated;—thus *gypsum*, *selenite*, or sulphate of calcium is found in transparent crystals, unchangeable in the air, but at a red heat crumbling down into the white powder known as plaster of Paris.

* The primitive forms which crystals assume may be arranged in six classes or systems, and in each of these there is a great number of secondary forms, for full details of which the student must refer to complete treatises on Mineralogy.

28. Some salts part with a portion of their water of crystallization when exposed to dry air, and crumble into powder, or, as it is termed, *effloresce;* as carbonate of sodium, acetate of calcium, sulphate of copper, &c.

29. Other salts, instead of parting with their water of crystallization, attract moisture from the air, and become more or less liquid or *deliquesce*, as, for instance, carbonate of potassium, chloride of calcium, &c.

EXERCISES ON CHAPTER XXVII.

1. What are oxygen-acids?
2. What are hydr-acids?
3. What are basic oxides?
4. What are salts?
5. How may salts be formed?
6. Why have some salts been termed Haloid salts?
7. What is shewn by the terminations *-ate* and *-ite?*
8. How are acids now regarded?
9. What is the modern theory of the formation of a salt?

10., 11., 12., 13., 14. Explain the action of *mono basic, di basic*, and *tri-basic* acids, or metals of different equivalent power.

15. How is an *acid-salt* or *super-salt* obtained?
16. Explain the terms *base* and *salt radical.*
17. Give examples of super-salts.
18. What is shewn by the prefixes *proto-, per-*, &c.?
19. What is a *basic* or *sub-salt?*
20. Explain and illustrate the term *double-salt.*
21. Why may some *super-* or *acid-*salts be considered as double salts?
22. What is meant by crystallization?
23. Under what circumstances do bodies crystallize? Give examples.
24. What is meant by *dimorphous* bodies?
25. When are bodies termed *isomorphous?*
26. What is meant by water of crystallization?
27. Give examples of *anhydrous* salts.
28. What are *efflorescent* salts?
29. When are salts said to be *deliquescent?*

CHAPTER XXVIII.

CARBONATES.

Carbonic acid H_2CO_3.

The Carbonates are formed by the union of carbonic acid with a metallic oxide, and are, for the most part, insoluble, white, or light-coloured salts. Many of them are found native, and nearly all may be artificially prepared by adding carbonic acid or a solution of an alkaline carbonate to a solution of a metallic salt. In the latter case a double decomposition ensues with interchange of acids and bases.

2. The affinity of carbonic acid for the majority of metallic oxides is so slight, that the addition of almost any other acid will affect its separation; the decomposition being accompanied with effervescence owing to the escape of carbonic acid gas, as in the familiar instance of the effervescing draughts prepared by adding tartaric acid to carbonate of sodium.

3. All the carbonates, excepting those of the alkalies, may be decomposed by heat. The alkaline carbonates are soluble, and retain a part, at least, of their acid at the highest temperatures.

4. **Carbonate of Potassium**, $K_2CO_3 + H_2O$, is prepared from the ashes of land plants, (especially of their green and tender parts), by stirring them in water so as to dissolve out all the soluble matter they contain. The *lye* thus obtained is strained, and then evaporated to dryness; the mass which remains is an impure carbonate of potassium (*grey salts*). This, when further purified, becomes the pearlash of commerce (*white salts*).

5. Pure carbonate of potassium may be obtained by heating cream of tartar (acid tartrate of potassium), $KH,C_4H_4O_6$, to redness, whereby the salt is decomposed, giving off inflammable gas and leaving behind a mixture of charcoal and carbonate of potassium ; the latter is dissolved out by water, filtered off, evaporated, and crystallized.

6. Potass being a powerful base is not entirely neutralized by carbonic acid, so that carbonate of potassium possesses alkaline properties, though to a less extent than the pure alkali.

Carbonate of potassium is deliquescent, insoluble in alcohol, but completely soluble in its own weight of cold water. When its crystals are heated, they lose their water of crystallization, and become anhydrous; they may be even fused without further change.

7. **Bicarbonate of Potassium**, $KHCO_3$, is formed by passing a stream of carbonic acid through a cold solution of the carbonate. It possesses very slight alkaline properties, and is much less soluble than the preceding

When a solution of it is boiled, or its crystals are heated, the salt is decomposed into water, carbonic acid, and the simple carbonate.

8. **Carbonate of Sodium**, $Na_2CO_3 + 10H_2O$, may be obtained from the ashes of sea-weeds, and marine plants, such as the *Salsola soda*, by a process similar to that described for preparing carbonate of potassium. The impure salt thus obtained forms the *barilla* of commerce, which is largely imported for soap-making, &c., from the coasts of Spain.

Sometimes the carbonate of sodium is found native, having been left by evaporation on the shores of certain lakes, termed *natron lakes*, in Africa ; also in Mexico, Hungary, &c. In some instances this *natron* is a sesquicarbonate of sodium, $2Na_2CO_3, H_2CO_3 + 3H_2O$.*

* The sesquicarbonate of sodium, like that of potassium, is a natural product, and cannot be formed at pleasure.

9. By far the greater part of the carbonate of sodium used at the present time is prepared from chloride of sodium (common salt), by converting it into the sulphate of sodium, and then decomposing the latter by carbonate of calcium (chalk).

The chloride of sodium is converted into sulphate of sodium by heating it with sulphuric acid as shewn in the preparation of hydrochloric acid, Chap. XVI., § 14.

This sulphate of sodium is then mixed with chalk and coal, and strongly heated in a reverberatory furnace, whereby, in the first place, the sulphate of sodium becomes reduced to sulphuret of sodium through the removal of its oxygen by the coal; and secondly, the sulphuret of sodium is decomposed by the carbonate of calcium with an interchange of elements, yielding sulphuret of calcium and carbonate of sodium. The latter is then dissolved out from the fused mass, dried, redissolved, purified and crystallized.

10. These three successive steps in the manufacture of carbonate of sodium may be shewn as follows:

1st. The production of sulphate of sodium from chloride of sodium.

$$2NaCl + H_2SO_4 = 2HCl + Na_2SO_4.$$

2nd. The conversion of sulphate of sodium into sulphuret of sodium.

$$Na_2SO_4 + 2C = 2CO_2 + Na_2S.$$

3rd. The double decomposition of sulphuret of sodium and carbonate of calcium.

$$Na_2S + CaCO_3 = CaS + Na_2CO_3.$$

11. Carbonate of sodium dissolves in about its weight of boiling water, and in twice its weight of c… the solution, like the corresponding solution of carbo… of potassium, is strongly alkaline, caustic, and deterge…

Crystals of carbonate of sodium *effloresce* in dry air, when heated, lose their water of crystallization, and … without further change.

12. **Bicarbonate of Sodium**, $NaHCO_3$ is a crystall…

* *Detergent*, cleansing, from *detergere* to wipe off.

white powder, possessing properties analogous to those of bicarbonate of potassium, and may be procured in a similar manner.

It cannot be dissolved in warm water without partial decomposition.

13. The carbonate of both potassium and sodium are much employed in glass-making, soap-making, washing, bleaching, and in medicine.

The carbonate of *sodium* forms a *hard* soap, that of *potassium*, a *soft* soap, when boiled with fatty matters.

The carbonate of sodium is rather less soluble than carbonate of potassium, but is more fusible, it consequently makes a better glass.

14. There appear to be at least three compounds of carbonic acid and ammonia.

1. The neutral **Carbonate of Ammonia** $(NH_3)_2H_2O,CO_2$, a white powder, formed by mixing one volume of carbonic acid gas with two volumes of gaseous ammonia.
2. **Sesqui-carbonate of Ammonia**, or ordinary "smelling salts," $2\{(NH_3)_2H_2O\}3CO_2$.—This substance may be readily prepared by mixing coarsely powdered sal-ammoniac (chloride of ammonium) with carbonate of lime, and exposing it to heat; mutual decomposition ensues, and sesqui-carbonate of ammonia sublimes, leaving a residue of chloride of calcium and lime.

$$4(NH_4)Cl+3CaCo_3=2\{(NH_3)_2H_2O\}3CO_2+2CaCl_2+CaO.$$

3. **Bicarbonate of Ammonia**, $(NH_3)H_2O,CO_2$, may be formed by passing a current of carbonic acid gas through dilute liquid ammonia. It is also the result of exposing the carbonate and sesqui-carbonate to the air, whereby part of the ammonia is dissipated.

The symbols in the preceding paragraphs refer to the carbonates of ammonia as salts of the volatile alkali, $(NH_3)_2H_2O$; they may, however, be expressed as salts of (the supposed mono-atomic metal) ammonium, NH_4, and will then appear in the following form:

Carbonate	of Ammonium	$(NH_4)_2CO_3$.
Sesquicarbonate	,,	$2\{(NH_4)_2CO_3\}CO_2$.
Bicarbonate	,,	$(NH_4)HCO_3$.

"A strong solution of sesquicarbonate of ammonium, made with tepid water, and filtered, warm, into a close vessel, will after a few days deposit crystals of the bicarbonate permanent in the air, and strongly resembling crystals of bicarbonate of potassium."—*Fownes.*

When organic substances containing nitrogen, as, flesh, animal excretions, &c., decay or putrefy, carbonic acid gas and gaseous ammonia, are set free; these uniting, escape into the air, as carbonate of ammonium, the powerful odor of which may be readily perceived near stable yards, manure heaps, &c.

15. **Carbonate of Barium, $BaCO_3$, is a** natural product, called *Witherite*, by mineralogists. It is highly poisonous, has no action on vegetable colors, and is but slightly soluble in water.

Salts of Barium are poisonous, those of Strontium are not.

16. **Carbonate of calcium, $CaCO_3$, is** distributed all over the globe, in the form of limestone, marble, chalk, coral-reefs, shells, &c.

It has been stated, that all the deposits of chalk and limestone, although they are so extensive as often to form mountain ranges, are of animal origin, and have been formed by accumulations of corals, shells, and skeletons of minute animals which formerly existed, but have become consolidated by pressure, and other causes, during the course of ages.

The formation of coral reefs, which are masses of carbonate of lime built up by myriads of coral-polyps, is an existing illustration of similar deposits.

17. Carbonate of calcium not unfrequently occurs [crys]tallized, as in *Iceland*, or *calcareous spar*, and in

mineral termed *arragonite*. These two minerals, though alike in composition, have very different crystalline forms, as well as different specific gravities, and optical properties. They are, in fact, illustrations of the dimorphous character of carbonate of calcium (See *Chap. XXVII.* § 24).

18. Carbonate of calcium is nearly insoluble in pure water, but is readily dissolved by water containing carbonic acid. Water so impregnated is termed *hard* water (page 40), and deposits carbonate of calcium when heated, owing to the escape of carbonic acid.

When the hardness of water arises from the presence of carbonate of calcium, it may be removed by the addition of potass, soda, or lime water, whereby the excess of carbonic acid is neutralized, and carbonate of calcium deposited. This is *one* use of potass, or soda, in washing.

19. The deposition of carbonate of calcium, from its solution in carbonated water, by evaporation, is the cause of the encrustations, found in tea-kettles and steam-boilers, especially in chalk and limestone districts; and also of the masses of mineral matter, often found in limestone caverns, termed *stalactites* and *stalagmites;* the former, hanging like icicles from the roof, the latter, rising in various shapes from the floor.

Ex. a.—Pass a current of carbonic acid through water which has been rendered milky by a little powdered chalk, the liquid will gradually become clear; if heated over a lamp, chalk will be again deposited.

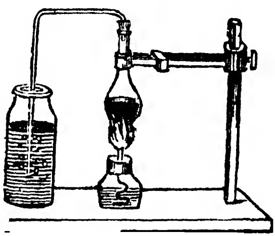


Fig. 75.

20. **Carbonate of Magnesium,** $MgCO_3$, is not easily prepared artificially; it occurs native in crystals somewhat resembling Iceland spar, but much more abundantly in combination with carbonate of calcium, forming ...esian limestone: $MgCa2CO_3$.

21. The carbonate of magnesium, or *magnesia alba* of the druggist, is not a neutral carbonate, but a compound of hydrate and carbonate, consisting of $4MgCO_3+MgH_2O_2+5H_2O$. It is prepared by adding a hot solution of carbonate of potass or soda to a hot solution of any soluble salt of magnesia, boiling the mixture for a few minutes, and then well washing and drying the precipitate formed.

22. **Carbonate of Iron,** $FeCO_3$, occurs in nature as *spathose iron ore*, and as *clay iron-stone*, the most abundant and important of the ores of iron. It is also often found in mineral waters, being held in solution by the excess of carbonic acid they contain; such waters are termed *chalybeate*, and may be known by the rusty color they impart to the stones, &c., over which they flow, and by their disagreeable, inky taste.

23. **Carbonate of Copper,** $CuCO_3$; a neutral carbonate of copper is unknown. When a *hot* solution of a salt of copper, (the sulphate, for instance,) is added to a *hot* solution of carbonate of potassium, or sodium, a bulky green precipitate is thrown down, which is a **Hydrated Carbonate of Copper,** $CuCO_3CuH_2O_2$, used by painters under the name of *green verditer*, or *mineral green.* The beautiful mineral called *malachite*, which is found chiefly among the Uralian mountains, is a native di-carbonate of copper.

24. If cold solutions be used instead of hot, the resulting precipitate will be of a blue colour; it is also a **Hydrated Carbonate of Copper,** $2(CuCO_3),CuH_2O_2$, known as *blue verditer*, and found beautifully crystallized in France, Saxony, Bohemia, &c., when it is termed *chessylite.*

25. **Carbonate of Lead,** $PbCO_3$, is a soft, white powder, insoluble in water; and called by various names, such as *ceruse*, *flake white, white lead*, &c. When ground with linseed, or some other oil, it is used as a white paint, or as the basis with which other paints are mixed in order to render them *opaque*, or, as it is termed, give them

If *pure* water be allowed to stand in leaden vessels, exposed to the air, it will, in a short time, contain particles of carbonate of lead, as stated in *Chap.* xxiiii., § 21.

26. White lead may be prepared by various processes, one of these (the Austrian method), is by passing a current of carbonic acid gas through a solution of *tris-acetate* or subacetate of lead, $Pb(C_2H_3O_2)_2,2PbO$, formed by boiling finely powdered *litharge* (protoxide of lead) in a solution of the acetate of lead, or in distilled vinegar (acetic acid, $H,(C_2H_3O_2)$. The subacetate of lead is decomposed by the carbonic acid, and carbonate of lead precipitated.

27. The oldest (the Dutch) process, and the one still mostly preferred, is by exposing thin plates of *pure cast* lead to the action of the vapor of vinegar, in earthen vessels, placed on a hot bed formed of manure or tanners' spent bark.

The vapor of vinegar or acetic acid appears in the latter process to act, chiefly, as a carrier between the carbonic acid, arising from the decomposition of the bark, &c., and the oxide of lead, formed under the influence of the acid vapor. Successive portions of the metal become converted into subacetate of lead, $Pb(C_2H_3O_2)_2,2PbO$; this subacetate is immediately decomposed by the carbonic acid, and carbonate of lead is formed. The process gradually proceeds from the exterior, till the whole thickness of the metallic lead is changed into a carbonate.

Ex. b.—Fill a series of test glasses with the following solutions, and add to each a few drops of a solution of carbonate of potassium, or sodium.

Solutions.		*Precipitates.*	
Nitrate of Barium	gives	Carbonate of Barium	*white.*
,, ,, Strontium	,,	,, ,, Strontium	,,
Chloride of Calcium	,,	,, ,, Calcium.	,,
Sulphate of Iron	,,	,, ,, Iron	,,
Chloride of Nickel	,,	,, ,, Nickel	*green.*
,, ,, Cobalt	,,	,, ,, Cobalt	*pink.*
Acetate of Lead	,,	,, ,, Lead	*white.*
Sulphate of Copper	,,	,, ,, Copper	*blue.*

EXERCISES ON CHAPTER XXVIII.

1. What are the carbonates, and how may they be prepared?
2. State how the carbonic acid may be separated from a carbonate, and give illustrations.
3. How are the carbonates affected by heat and moisture?
4. How is ordinary carbonate of potassium obtained?
5. How may carbonate of potassium be obtained pure?
6. What are the properties of carbonate of potassium?
7. How is bicarbonate of potassium procured, and in what does it differ from the carbonate?
8. Whence may carbonate of sodium be obtained?
9. How is carbonate of sodium manufactured on a large scale?
10. Explain the successive steps in the usual process of manufacturing carbonate of sodium.
11. State the properties of carbonate of soda.
12. How is the bicarbonate of sodium obtained, and what are its properties?
13. For what are the carbonates of potassium and sodium used?
14. Describe the carbonates of ammonium.
15. What are the properties of carbonate of barium?
16. Where is carbonate of calcium found?
17. Illustrate the dimorphous character of carbonate of calcium.
18. What is *hard* water?
19. Explain the origin of *stalactites*, &c.
20. Where is carbonate of magnesium found?
21. What is *magnesia alba?*
22. Where is carbonate of iron found?
23. What is *green verditer*, and how may it be prepared? What is *malachite*, and where is it chiefly found?
24. What is *blue verditer?*
25. What is carbonate of lead, and for what is it used?
26. How is white lead prepared by the Austrian process?
27. Explain the common process for obtaining white lead.

CHAPTER XXIX.

SULPHATES.

Sulphuric Acid = H_2SO_4.

1. **The Sulphates** form an important class of salts. They are produced, either by the direct combination of sulphuric acid with a metallic oxide,—by the addition of sulphuric acid to a metallic carbonate, whereby the carbonic acid is displaced,—or by the action of sulphuric acid on a metal.

In the latter case the metal is oxidized, either at the expense of the water with which the acid is diluted, while hydrogen gas is evolved, see *Chap.* v. § 13, 14, or at the expense of a portion of the acid, which, *if concentrated*, is decomposed, and sulphurous acid gas given off. See *Chap.* xiv. § 12.

2. The sulphates are sometimes formed in nature, through the oxidizing influence of the air on metallic sulphurets, whereby the metal is converted into an oxide and the sulphur into sulphuric acid, and these uniting together form the sulphate. Thus, for instance,

$$FeS + 4O = FeSO_4.$$

3. The sulphates are for the most part soluble in water,* and the solutions give a white precipitate with chloride of barium, BaCl, or any other soluble barytic salt; this precipitate (sulphate of barium, $BaSO_4$) is insoluble, either in acids or alkalies.

* There are, so far as at present known, but six *insoluble* sulphates, viz., sulphate of barium, bismuth, tin, lead, antimony, and mercury; and six *sparingly soluble* ones, viz., sulphate of strontium, lime, silver, zirconium, yttrium, and cerium.

Chloride of barium is thus a test for the presence of sulphuric acid.

$$BaCl_2 + H_2SO_4 = 2HCl + BaSO_4.$$

4. An insoluble sulphate may be detected by being fused with three times its weight of carbonate of potassium or sodium. Double decomposition ensues, a metallic carbonate and an alkaline sulphate being formed. The nature of the metal may be ascertained from the former, and the amount of sulphuric acid from the latter.

5. All the sulphates, except those of potassium, sodium, lithium, barium, strontium, and calcium are decomposed by a white heat, yielding sulphuric acid, sulphurous acid, and oxygen, and leaving a metallic oxide.

6. All the sulphates are decomposed when mixed with charcoal and heated to redness, being for the most part converted into sulphurets, while carbonic acid and carbonic oxide are given off.

7. **Sulphate of Potassium**, K_2SO_4, and **Acid-sulphate of Potassium**, $KHSO_4$, are artificial salts. The latter formerly termed *sal-enixum*, is formed during the process of preparing nitric acid from nitrate of potassium by the action of sulphuric acid, *Chap. VIII.*, § 21. It is used in calico-printing, and for cleaning brass and other metallic work previous to lacquering or finishing, also for cleaning coins. When heated it is decomposed into sulphuric acid and the neutral sulphate.

The sulphate of potassium may be obtained by adding potass or carbonate of potassium to a solution of the acid-sulphate; it is sometimes used in medicine.

8. **Sulphate of Sodium**, or *Glauber's Salts*, $Na_2SO_4 + 10H_2O$, may be prepared by adding sulphuric acid to common salt, and is formed in large quantities in the processes for ob-

taining chlorine and hydrochloric acid, described in *Chap. XVI.*, §§ 9, 14, and in the manufacture of carbonate of sodium, *Chap. XXVIII.* § 10.

$$\text{(i.) } 2NaCl + MnO_2 + 2(H_2SO_4) = 2Cl + 2H_2O + MnSO_4 + Na_2SO_4.$$

$$\text{(ii.) } 2NaCl + H_2SO_4 = 2HCl + Na_2SO_4.$$

It is termed Glauber's salt from a chemist named Glauber, who first introduced it as a medicine; it is now chiefly used for horses and cattle.

Ex. a.—Put half an ounce of sulphate of sodium, with the same quantity of nitrate of ammonium, into a mortar, and let them be smartly rubbed together; they will each part with their water of crystallization, and this water will render them both quite fluid.

9. **Sulphate of Barium,** $BaSO_4$, is an abundant natural product, known as *heavy spar.* Its artificial production has been noticed in § 3. It forms a valuable white paint (*permanent white*), which remains untarnished by sulphuretted hydrogen or other vapors. It is frequently employed for the purpose of adulterating white lead; the adulteration may be detected by dissolving white lead in nitric acid; if pure, the whole will be dissolved; if any sulphate of barium be present, it will remain as a white sediment.

10. **Sulphate of Calcium,** or *Gypsum,* $CaSO_4 + 2H_2O$, is a soft, white, mineral, very abundant in nature; the finer kinds are known as *alabaster,* and when regularly crystallized, it is called *selenite.* When gypsum is heated to about 300° F., the greater part of its water of crystallization is expelled; it is then ground to powder, and known as *Plaster of Paris.*

11. When Plaster of Paris is made into a thin paste with water, it again combines with it, and speedily *sets* or hardens, as we see in the case of plaster busts, stucco work, &c.

This effect, however, does not take place if the gypsum

has been over-heated, so as to become anhydrous, or if the plaster has been made some time.

12. Since sulphate of calcium is somewhat soluble, and exists abundantly in the earth, it is frequently found in spring or well water. All waters containing lime are called *hard*, but such as contain sulphate of calcium are *permanently* hard, that is, they do not deposit their lime and become soft by boiling, as waters containing only carbonate of calcium do; such water may, however, be deprived of its hardness by the addition of an alkaline carbonate, which, by double decomposition, precipitates the lime in the form of chalk.

13. Ground gypsum is frequently used in agriculture, as a manure for clover, peas, beans, &c. It is believed to act by furnishing lime and sulphur to the plants, and by absorbing the carbonate of ammonium found in the air and in rain-water. It is said to *fix* the ammonia; this it does by double decomposition, whereby the volatile carbonate of ammonium becomes the non-volatile sulphate; and the gypsum or sulphate of calcium becomes carbonate of calcium or chalk.

$$CaSO_4 + (NH_4)_2CO_3 = (NH_4)\ SO_4 + CaCO_3.$$

Ex. b.—Put into a glass a few spoonfuls of a saturated solution of chloride of calcium, and add to it gradually, drop by drop, sulphuric acid. If these two liquids be stirred together with a glass rod, they become converted into an opaque, white, and almost solid mass of sulphate of calcium.

14. **Sulphate of Magnesium**, $MgSO_4 + 7H_2O$, is well known by the name of *Epsom salts*, having been first obtained from a spring at that place. It is found in sea-water, and in many mineral springs. It has a bitter, nauseous taste, is very soluble, crystallizes readily, *effloresces* slightly when exposed to the air, and is much used as an aperient medicine.

15. It is manufactured in large quantities by the action

of dilute sulphuric acid on magnesian lime stone (a native carbonate of calcium and magnesium). Sulphates of calcium and of magnesium are formed, but the latter being very soluble, is easily separated from the slightly soluble sulphate of calcium, by adding a little water to the mixture, and filtering.

16. Sulphate of magnesium is also frequently obtained by the action of sulphuric acid on *bittern*, that is, the liquid which remains after the extraction of the common salt from sea-water.

This *bittern*, so called from its bitter taste, consists chiefly of sulphate of magnesium and chloride of magnesium. On being boiled with sulphuric acid, the chloride of magnesium is decomposed, and sulphate of magnesium formed.

The latter process is preferred, when the sulphate of magnesium is prepared for the purpose of being converted into the carbonate, on account of the absence of iron, which metal is generally present in the salt prepared from magnesian limestone.

On the other hand, the former process, § 17, is preferred for medical purposes, because the salt is free from chloride of magnesium, and consequently not *deliquescent*, and may be obtained very nearly pure.

17. **Sulphate**

$18H_2O$, may be prepared by boiling *calcined* (or burnt) clay in sulphuric acid; it is a very soluble salt, which is unaltered in the air; when exposed to heat, it loses its water of crystallization and becomes *anhydrous*; heated to redness, it is decomposed, the acid being expelled, while pure anhydrous alumina remains.

Sulphate of aluminum is itself an unimportant compound, but combines with the sulphates of potassium, sodium, and ammonium, forming double salts of great interest, known as *alums*.

18. **Sulphate of Aluminum and Potassium,** *Common Alum*, $Al_2K_24SO_4+24H_2O$, may be prepared by simply mixing together the sulphates of aluminum and potassium, but it is generally obtained from a kind of slaty clay or *shale*, which contains much *iron-pyrites*, bisulphuret of iron, FeS_2, by the following process.

The aluminous shale, or *alum-schist*, as it is termed, is moderately roasted, then moistened and exposed to the action of the air, whereby the sulphur becomes oxidized into sulphuric acid, which, uniting with the iron and alumina, forms proto-sulphate of iron and sulphate of aluminum.

These salts are dissolved out from the mass by water, the solution partly evaporated, and left to cool, when great part of the sulphate of iron separates in crystals, and the sulphate of aluminum being much more soluble, remains. Chloride of potassium is then added, whereby the rest of the sulphate of iron is decomposed, while sulphate of potassium and chloride of iron are formed. The sulphates of aluminum and potassium then combine and form alum, which is separated by crystallization, the more soluble chloride of iron remaining in solution.

19. These steps may be thus stated:—

1st. Oxidizing influence of the air,

$$Al_2O_3 + 3FeS_2 + 21O = Al_23SO_4 + 3FeSO_4.$$

2nd. Separation of a part of the sulphate of iron by crystallization, decomposition of the remainder by chloride of potassium, and subsequent formation of the double salt.

$$Al_23SO_4 + 3FeSO_4 + 2KCl = (Al_23SO_4 + K_2SO_4) + FeCl_2.$$
$$= Al_2K_24SO_4 + FeCl_2.$$

20. Alum has a sweetish, astringent taste, reddens vegetable blues, is soluble in eighteen times its weight of cold water (60° F.), and in about three-fourths of its weight of boiling water. When heated, it becomes anhydrous (*burnt alum*); at a high temperature it is decomposed.

Alum is much used in tanning, dyeing, calico printing, &c.; and by candle makers, to harden the tallow and render it white.

The pigments called lake-colors, such as crimson-lake, madder-lake, carmine, &c., are formed by adding alum to a solution of the coloring matter, and then an alkali to precipitate the alumina.

Ex. b.—Repeat *Ex. a*, *Chap.* xxi.

21. Alum affords a striking example of *isomorphism.* (*Chap. XXVII.*, § 25); for the potass may be replaced by soda, or ammonia; and the sesquioxide of aluminum by sesquioxide of chromium or iron without altering the crystalline form of the salt.

Common alum	$= Al_2K_24SO_4$,	*or*	$Al_23SO_4 + K_2SO_4$.
Soda alum	$= Al_2Na_24SO_4$,	*or*	$Al_23SO_4 + Na_2SO_4$.
Ammonia alum	$= Al_2(NH_4)_24SO_4$,	*or*	$Al_23SO_4 + (NH_4)_2SO_4$.
*Chrome alum	$= Cr_2K_24SO_4$,	*or*	$Cr_23SO_4 + K_2SO_4$.
*Iron alum	$= Fe_2K_24SO_4$,	*or*	$Fe_23SO_4 + K_2SO_4$.
*Manganese alum	$= Mn_2K_24SO_4$	*or*	$Mn_23SO_4 + K_2SO_4$.

A very pure alum (*Roman alum*) is obtained in the Roman States, by simply heating a mineral termed *alum-stone*, till sulphurous acid begins to escape from it, and then dissolving the residue in water.

Alum-stone contains an insoluble subsulphate of alumina with sulphate of potass. The heating has the effect of separating the excess of alumina, so that a neutral sulphate of alumina is formed; which, uniting with the sulphate of potass, forms alum. Alum-stone appears to be continually produced at the Solfatara, near Naples, and in other volcanic districts, by the joint action of sulphurous acid and oxygen upon *trachyte*, a volcanic rock composed almost entirely o. *felspar*, which is a double silicate of alumina and potass—

$$Al_43SiO_4 + K_4SiO_4 = 4(AlKSiO_4).$$

22. **Proto-sulphate** or **Sulphate of Iron,** *Copperas*, or *Green vitriol*, $FeSO_4 + 7H_2O$.—This very common, but important salt is obtained on a large scale by exposing *iron-pyrites*, bisulphuret of iron, FeS_2, to the action of air and moisture. By absorbing oxygen, the sulphur becomes sulphuric acid, and the iron, protoxide of iron,—these combined, form protosulphate, or, as it is more simply termed, sulphate of iron. As, however, there is a great excess of sulphuric acid formed, pieces of old iron are added to saturate or neutralize it, and the salt is obtained in crystals, on evaporation.

* These substances are termed *alums*, although they contain no alumina, because they are obtained from alum by substitution.

23. Sulphate of iron is soluble in about twice its weight of cold water, and has a disagreeable, inky taste; its crystals, which are of a green color, effloresce in dry air, but oxidize slowly when exposed to moisture, being changed on the surface into *persulphate of iron*, and becoming of a rusty, reddish color.*

24. Sulphate of iron, when moderately heated, loses six sevenths of its water of crystallization, and becomes white; if exposed to a greater heat, it becomes nearly anhydrous, and when heated to redness it is decomposed, as shewn in the process for obtaining Nordhausen, or fuming sulphuric acid, *Chap.* xiv., § 18—

$$4FeSO_4 + H_2O = 2Fe_2O_3 + 2SO_2 + H_2O,2SO_3.$$

Ex. c.—Drop some crystals of protosulphate of iron into *strong* sulphuric acid, or alcohol, they will not be dissolved, but be converted into a white powder (anhydrous sulphate of iron), the liquid having deprived the salt of its water of crystallization.

25. **Persulphate or sesqui-sulphate of Iron**, $Fe_2 3SO_4$; this substance, which is a sulphate of the peroxide of iron, is obtained by oxidizing the protosulphate, as shewn in the last two paragraphs; or, more easily, by adding to a solution of the common sulphate, as much sulphuric acid as it already contains, heating the mixture to the boiling point, and then gradually dropping in nitric acid, so long as it darkens the solution. On evaporation, a reddish brown, uncrystallizable mass is obtained, which slowly dissolves in water.

Persulphate of iron combines with sulphate of potass, soda, or ammonia, forming a substance already noticed as iron-alum, similar to common alum, in taste, composition, and form of crystals.

26. Common sulphate of iron is much used for dyeing dark colors, for making ink, and as a tonic in medicine.

* "Whence the French term *couperose* applied to them, corrupted into *copperas*."—*Brande.*

Ex. d.—To a solution of sulphate of iron, add some solution of nutgalls (or a few drops of the tincture of galls), a bluish-black precipitate is formed, a tanno-gallate of the protoxide of iron. If this precipitate be held in suspension in the water by means of a little gum dissolved in it, we have common ink.

The color darkens by degrees, in consequence of the peroxidizement of the iron, by the action of the air But ink affords a more durable writing when used in the pale state, because its particles are then finer, and penetrate the paper more intimately. When ink consists chiefly of tanno-gallate of *peroxide* of iron, however black, it is merely superficial, and is easily erased or effaced.

Ex. e—Add some hydrochloric acid to the black solution just made, the color will disappear; soluble chloride of iron is formed.

Ex. f.—Instead of solution of galls, in the last experiment, add a little solution of ferro-cyanide of potassium, a blue precipitate will be formed (Prussian blue); deep blue, if *per*sulphate or any *per*salt of iron be used, and pale blue, darkening on exposure to the air, if *proto*sulphate or any *proto*salt of iron be employed.

27. **Sulphate of Copper,** $CuSO_4$, the common *blue vitriol* of commerce, is much used in the arts, for precipitation by the electrotype process, and as the source of numerous paints and dyes, and in surgery as an external caustic application.

28. Sulphate of copper may be prepared by dissolving oxide of copper in sulphuric acid (*Chap.* xiv. §§ 12, 13); it is obtained in considerable quantities, and very pure, in certain processes for refining gold and silver, termed *parting*; and is manufactured on a large scale by exposing sulphuret of copper, *copper pyrites*, CuS, to the action of air and moisture, the latter process being analogous to that already described, § 22, for preparing sulphate of iron.

Ex. g.—Make a very weak solution of sulphate of copper, and add to it liquid ammonia; it will become of a most beautiful blue color, such as we see in the shop windows of the chemists.

29. **Sulphate of Zinc,** $ZnSO_4$; this substance, formerly called *white vitriol,* is prepared on a large scale by roasting and oxidizing the sulphuret of zinc, *zinc blende,* ZnS; it is also the residue formed in the usual process for

obtaining hydrogen gas, *Chap*. v. § 13. It is very similar in appearance to sulphate of magnesium, and is not unfrequently used in medicine as an emetic, and as a cooling application in inflammation of the eyes. Like all the other salts of zinc, it is poisonous.

EXERCISES ON CHAPTER XXIX.

1. What are sulphates, and how may they be formed?
2. How are sulphates naturally formed?
3. How may a soluble sulphate be recognised?
4. How may an insoluble sulphate be analysed?
5. How are the sulphates affected by heat?
6. What effect is produced by heating the sulphates with charcoal?
7. Describe the preparation and uses of sulphate of potass.
8. How is sulphate of sodium prepared, and for what is it used?
9. What is sulphate of barium, and for what is it used?
10. What are *alabaster*, *gypsum*, *selenite*, and plaster of Paris?
11. What is the cause of plaster of Paris *setting* or hardening?
12. What is meant by water being *permanently hard?*
13. Of what use is sulphate of calcium in agriculture?
14. State the composition and properties of Epsom salts.
15. How is sulphate of magnesium generally prepared?
16. How may sulphate of magnesium be obtained from sea-water?
17. How is sulphate of aluminum formed?
18. What is common alum, and whence is it obtained?
19. Describe the process of the alum manufacture.
20. State the properties of alum.
20. For what is alum used?
21. Describe the different forms of alum.
22. How is the common or *proto*-sulphate of iron obtained?
23. State the properties of sulphate of iron.
24. How is sulphate of iron affected by heat?
25. What is persulphate of iron?
26. For what is sulphate of iron used?
27. For what is sulphate of copper employed?
28. How is sulphate of copper obtained?
29. State the mode of preparation and the uses of sulphate of zinc.

CHAPTER XXX.

NITRATES.

Nitric acid = HNO_3.

1. **The Nitrates** may be prepared by the direct action of nitric acid on the metals, metallic oxides, or metallic carbonates; in the first case with decomposition of the acid, (see *Chap.* viii. § 14); and in the last case with evolution of carbonic acid.

2. All the nitrates are soluble in water, and they are all decomposed by heat and by sulphuric acid.

The only native nitrates known are those of potassium, sodium, calcium, and magnesium.

3. **Nitrate of Potassium,** *saltpetre*, KNO_3; this valuable salt is formed naturally in the soil and appears as a sort of efflorescence on its surface, in various hot and dry countries, and especially in the East Indies, whence most of the saltpetre used in Great Britain is obtained.

4. In France and other parts of the continent it is produced artificially in what are termed *nitre-beds*. These are heaps of vegetable and animal refuse, mixed with old mortar, plaster, rubbish, or other calcareous* matter, sheltered from the rain, but freely exposed for some time to the action of the air. In the gradual putrefaction of the animal substances which ensues, their nitrogen combines with oxygen and forms nitric acid, this unites with the lime and potass present in the heap, forming nitrates of calcium and potassium.

* *Calcareous*, containing lime.

The plaster of old walls frequently contains as much as 5 per cent of nitre, which appears to arise from the decomposition of the animal matter contained in the mortar.

The formation of nitric acid, above mentioned, takes place only in the presence of alkaline substances. When they are absent, the nitrogen passes off combined with hydrogen, forming ammonia, NH_3.

Even in the presence of lime there is reason to believe that ammonia is first formed, and its constituents afterwards converted into nitric acid and water, thus:

$$NH_3 + 4O = HNO_3 + H_2O.$$

5. The mixture of nitrates of calcium and potassium obtained from the *nitre-beds* is dissolved in water together with wood-ashes (carbonate of potassium), whereby the former is decomposed, carbonate of calcium being precipitated, while nitrate of potassium remains in solution.

$$Ca2NO_3 + K_2CO_3 = CaCO_3 + 2KNO_3.$$

This solution of nitrate of potassium, as well as that obtained by the process of *lixiviation** from naturally nitrous soils, is then drawn off, evaporated, and crystallized.

It may be further purified by re-solution and crystallization several times repeated.

6. Nitrate of potassium crystallizes in the form of six-sided prisms; is anhydrous, colorless, and inodorous; has a cooling saline taste; and is readily soluble both in hot and cold water. At about 616° it fuses without change, and cools in a white mass, sometimes cast into cakes or balls, (*sal prunella*). At a higher temperature it is decomposed affording abundance of oxygen gas.

"One pound of nitre yields about 12,000 cubic inches ot oxygen sufficiently pure for many common experiments; it is, of course, mixed with a portion of nitrogen."—*Brande.*

Ex. a.—The cold produced by saltpetre mixed with water is very great, so that if one part of powdered nitre be mixed with five of water, the thermometer which stood at 60° in the water, will sink to 45° or even lower, in the mixture.

* *Lixiviation* is the separation of soluble substances from a mass of matter, by pouring water over it, as in obtaining carbonate of potassium from wood ashes.

7. Saltpetre is powerfully *antiseptic,* that is, it preserves animal substances from putrefying, hence it is used with common salt, for salting meat, to which it imparts a reddish color. It is the chief source of nitric acid, which is obtained from it by the action of sulphuric acid, (see *Chap.* viii. § 21). In medicine it is used in small doses as a sedative, but in quantities of more than half an ounce it is poisonous. The principal use of saltpetre, however, is in the manufacture of gunpowder.

8. The large quantity of oxygen contained in nitrate of potassium, and the readiness with which this salt may be decomposed, render it most suitable for rapid and sudden combustion, hence it is the chief ingredient in gunpowder and fire works.

Sulphur, phosphorus, carbon, and most of the metals when in filings or powder, *deflagrate* (*i.e.* burn suddenly and vividly) if thrown on red hot nitre, sometimes with a considerable explosion.

Ex. b.—Mix together three parts of nitre, two of dry carbonate of potassium and one of sulphur. Put a few grains of this powder into a ladle, and place it over a fire, so that the powder may heat very gradually. It will, in a minute or so, turn black, fuse, emit a faint blue flame, and then explode with a tremendous report.

Ex. c.—Heat some nitre to redness in a crucible, throw a little sulphur into it, combustion of the sulphur will take place, nitrogen escape, and a mixture of sulphate and sulphite of potassium remain.

Ex. d.—Throw some shreds of phosphorus upon red hot nitre, instant inflammation will ensue, and a phosphate of potassium be formed while a large quantity of nitrogen escapes.

Ex. e.—A mixture of five grains of powdered nitrate of potassium with one grain of phosphorus, will produce a very violent explosion when struck on an anvil by a hot hammer. Nitrogen gas, phosphoric acid, and phosphate of potassium, are the results of this decomposition.

Ex. f.—If half an ounce of nitrate of potassium be heated in a crucible, and a quarter of an ounce of powdered charcoal be thrown into it, a brilliant combustion will take place. The new products are carbonate of potassium, which remains in the crucible, and carbonic acid and nitrogen gases, which escape.

9. **Gunpowder** is a mechanical mixture of saltpetre, charcoal, and sulphur, in different proportions, according to the purposes to which it is to be applied.

	Nitre.	*Sulphur.*	*Charc.*
Common powder	75	$12\frac{1}{2}$	$12\frac{1}{2}$
Shooting ,,	78	10	12
Miners' ,,	65	20	15

The larger the proportion of sulphur, the less powerful will be the explosion of the powder, and the less will it be injured by damp.

10. The explosive violence of gunpowder depends upon the sudden production of gaseous compounds, which are at the same time greatly expanded by the heat evolved. The gases given off are carbonic oxide, carbonic acid, nitrogen, and sulphurous acid, while the solid matter that remains, consists chiefly of sulphuret of potassium, with (occasionally) sulphate and carbonate of potassium and charcoal.

11. **Nitrate of Soda,** *cubic nitre,* $NaNO_3$; this substance is known in commerce as Chili saltpetre, large natural beds of it having been found in that country. It may be employed in many cases instead of nitrate of potassium.

Professor Proust says, "that five parts of nitrate of sodium, with one of charcoal, and one of sulphur, will burn three times as long as common powder, so as to form an economical composition for fireworks. It, however, gets damp by exposure to the air, which the compositions of saltpetre do not, if the salt be pure."

12. **Nitrate of Calcium,** $Ca2NO_3$: this substance abounds in old mortar and plaster walls, and is formed in the nitre-beds, as already mentioned. It deliquesces on exposure to the air. No use is made of it except in the manufacture of saltpetre, § 5.

13. **Nitrate of Ammonium,** $(NH_4)NO_3$, is a deliquescent salt, very soluble in water, prepared by neutralizing dilute nitric acid with carbonate of ammonium and evaporating the solution. When *carefully* heated to between

400° and 500°, it is decomposed into protoxide of nitrogen (laughing gas), and water. (See *Chap. VIII.*, § 11.)

$$N_2O+2H_2O.$$

At 600° it explodes with considerable violence.

14. **Nitrate of Silver,** *lunar caustic,* $AgNO_3$, is easily prepared by dissolving silver in dilute nitric acid; the solution, if pure, is perfectly transparent and colorless; white, anhydrous crystals, in the form of thin plates, may be obtained by evaporation.

15. Nitrate of silver fuses at 426°, and may be cast into small, cylindrical (pencil-like) sticks; in this form it is used by surgeons as an *external* application, to *cauterize* or burn away diseased flesh, warts, &c. At 700° it is decomposed into oxygen, nitrous acid and metallic silver.

Ex. g.—Throw a few crystals of nitrate of silver on red-hot charcoal, they will be decomposed, light and heat will be evolved, and the charcoal covered with spangles of metallic silver.

16. The nitrate (like the other salts of silver) is decomposed and blackened on exposure to light, and is no longer perfectly soluble in water. When in contact with organic substances it communicates to them, after a short time, a deep purple or black tint. This blackening is due to the separation of finely divided metallic silver, while the caustic effect of the salt is produced by the nitric acid.

17. This tendency of nitrate of silver to become black when exposed to light, occasions it to be much used for marking linen, for dyeing the hair, and especially for photographic purposes.

Ex. h.—Dissolve a little nitrate of silver in a wine-glass full of water, and immerse in it a slip of bright, clean, copper. Decomposition will take place, the copper unites with the nitric acid forming nitrate of copper, and the silver will be precipitated.

$$2AgNO_3+Cu = Cu2NO_3+2Ag.$$

Ex. i.—A still more beautiful effect takes place when several drops of a solution of nitrate of silver are let fall on a plate of polished cop-

per. Here, in a very short time, a brilliant precipitation of metallic silver will take place, in an arborescent or branching form.

Ex. k.—Immerse a piece of white silk ribbon in a solution of nitrate of silver, and while wet expose it to a stream of hydrogen gas; it will at first be blackened, and afterwards become silvered by the reduction of the metal. Ivory, marble, and other bodies may be stained black, or even silvered, by moistening them with nitrate of silver, and exposing them to the full power of the sun's rays. If two or three coatings be applied, and the articles polished, when dry, with a piece of leather, they will have the appearance of silver.

Ex. l.—Nitrate of silver, like the nitrates of potassium, &c., explodes with charcoal, sulphur, or phosphorus, when struck with a *hot* hammer. N.B. *Very small* quantities of the substances should be employed, and great caution exercised.

18. **Nitrates of Mercury.** There are several nitrates of mercury; they are of little use, and their true composition is not well ascertained. The **Protonitrate of Mercury,** (*nitrate of the protoxide*), $HgNO_3$, is valuable as a chemical test.

19. **Nitrate of Copper,** $Cu2NO_3$, is prepared by the action of nitric acid upon copper, (see *Chap. VIII.*, § 14.) Its crystals are of a deep blue color, very deliquescent and soluble; when heated to redness, they yield protoxide of copper, CuO.

Ex. m.—Wrap some crystals of nitrate of copper in tinfoil; while dry, no chemical action takes place, but moisten them with a drop of water, and soon heat will be evolved and sparks of fire flash out.

20. **Nitrate of Barium,** $Ba2NO_3$; and **Nitrate of Strontium,** $Sr2NO_3$.—These salts are not much used except in the manufacture of coloured fires. The salts of barium give a green, and the salts of strontium, a red colour to flame. Nitrate of barium is sometimes used as a test for sulphuric acid.

Ex. n.—*Red Fire.* Take 40 parts of dried nitrate of strontium, 13 of sulphur, 4 of sulphuret of antimony, and 5 of chlorate of potassium.

Ex. o.—*Green Fire.* Use 77 parts of dried nitrate of barium in the preceding, instead of nitrate of strontium.

EXERCISES ON CHAPTER XXX.

1. How may the nitrates be prepared?
2. State some general characteristics of the nitrates.
3. Where is nitrate of potassium found native?
4. What are *nitre beds?*
5. How is nitrate of potassium obtained from the nitre beds?
6. State some of the properties of nitrate of potassium.
7. Name some of the chief uses of nitrate of potassium.
8. Why is nitre adapted for fire-works, &c.?
9. What is gunpowder?
10. How may the explosive force of gunpowder be accounted for?
11. Whence is nitrate of sodium obtained?
12. How is nitrate of calcium obtained?
13. State some of the properties of nitrate of ammonium.
14. How is nitrate of silver obtained?
15. How is nitrate of silver affected by heat?
16. How is nitrate of silver affected by light and organic matter?
17. What are the chief uses of nitrate of silver?
18. Give the formula of protonitrate of mercury, and state its uses.
19. What is nitrate of copper?
20. For what are the nitrates of barium and strontium employed?

CHAPTER XXXI.

PHOSPHATES, CHLORATES, ETC.

1. Three different forms of phosphoric acid exist:—

Phosphoric Acid (*tribasic*) . . $= H_3PO_4$.
Pyrophosphoric Acid (*dibasic*) . $= H_4P_2O_7$.
Metaphosphoric „ (*monobasic*) $= HPO_3$.

The phosphates, with the exception of those of the alkalies, are but sparingly soluble in pure water; they are all dissolved, without effervescence, in an excess of phosphoric or nitric acid.

2. **Rhombic Phosphate of Sodium,** $Na_2HPO_4 + 12H_2O$, is a tribasic phosphate, having an alkaline reaction; it is occasionally used in medicine as a purgative (*sal perlatum*), and in chemistry as a re-agent for procuring the other phosphates.

3. If common or caustic soda be added to a solution of the rhombic phosphate of sodium, the atom of replaceable hydrogen in each atom of the phosphate will be replaced by sodium, and a neutral phosphate of sodium, Na_3PO_4, sometimes termed **Subphosphate of Sodium**, will be obtained.

4. If phosphoric acid be added to a solution of the common or rhombic phosphate of sodium, a **Superphosphate or Acid Phosphate of Sodium**, will be formed, NaH_2PO_4.

5. By strongly heating the common phosphate of sodium, *all* the water is expelled, and **Pyrophosphate of Sodium**, $Na_4P_2O_7$, remains; for $Na_4P_2O_7 = 2Na_2HPO_4—H_2O$. An acid pyrophosphate of sodium is also known, having the composition, $Na_2H_2P_2O_7$.

6. When either superphosphate of sodium, or microcosmic salt, is strongly heated, water, or water and ammonia are driven off, and a monobasic phosphate remains, termed the **Meta-phosphate of Sodium**, having the composition, $NaPO_3$.

Ex. a.—The tribasic phosphates of soda give a bright *yellow* precipitate when added to a solution of nitrate of silver. The monobasic and dibasic phosphates give a *white* one.

7. **Microcosmic Salt** is a double phosphate of sodium and ammonium, $Na(NH_4)HPO_4 + {}_4H_2O$, much employed as a *flux* in blowpipe experiments. Both the ammonia and water are driven off by heat, and a very soluble and fusible glassy substance (metaphosphate of sodium) is left.

8. The phosphates of sodium are instances of the formation of different substances by substitution or replacement, as in the following series of tribasic phosphates, in which either one, two, or three of the atoms of hydrogen in the acid, may be replaced by one, two, or three atoms of another metal: thus—

$$\left.\begin{array}{lll} \textbf{Subphosphate of Sodium} & & = Na,Na,Na, \\ \textbf{Rhombic Phosphate} & ,, & = Na,Na,H, \\ \textbf{Superphosphate} & ,, & = Na,H,H, \\ \textbf{Microcosmic Salt} & & = Na,(NH_4)H, \end{array}\right\} PO_4.$$

Similarly,—

$$\left.\begin{array}{l}\textbf{Pyrophosphate of Sodium, } (\alpha) = Na_2,Na_2, \\ \quad ,, \qquad ,, \qquad ,, \qquad (\beta) = Na_2,H_2,\end{array}\right\} P_2O_7.$$

9. **Phosphate of Calcium** exists in several forms corresponding to the phosphates of soda already named. The most important of these is the *bone-earth* phosphate, Ca_32PO_4, so termed from its being the principal mineral ingredient in bones (see *Chap. XV.* §§ 5, 6): it is obtained by heating the bones of animals till all the organic matter is dispelled.

10. Since the bones and flesh of animals contain various phosphates, these must be supplied by the food on which they live. They are, therefore, found in large quantities in the grain and seeds of plants. Every crop removes some of the phosphates from the soil, and will thus diminish its productiveness. Hence burnt bones are valuable as a manure, for they restore to land those earthy phosphates which are indispensable to first-rate crops of grain.

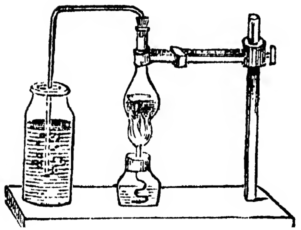
"Phosphate of calcium is essential to the healthful growth of corn; a soil may be apparently rich, and produce abundance of straw, but if phosphoric acid and lime be deficient, the wheat will be light and small."

11. Phosphate of calcium occurs native in the minerals *apatite* and *phosphorite.* Massive beds of it are found at Estremadura in Spain, and also in the United States. It forms the *burnt hartshorn* of the druggist, and is used for absorbing grease from cloth, for making *cupels,* (see *Chap. XXVI.* § 20), and for preparing phosphorus.

12. The **Chlorates**, $MClO_3$, very much resemble the nitrates; they deflagrate, however, with greater violence, when mixed with inflammable substances. They are decomposed by a red heat, nearly all of them being resolved into metallic chlorides, and evolving pure oxygen gas.

$$MClO_3 = MCl + 3O.$$

13. Chlorate of Potassium, $KClO_3$.—This interesting salt may be prepared by passing a current of chlorine gas through a solution either of potassium, or of carbonate of potassium, till the gas ceases to be absorbed. The liquid which contains both chlorate and chloride of potassium may be partially evaporated and allowed to cool, when crystals of chlorate of potassium will be deposited; by further evaporation and cooling, crystals of the more soluble chloride of potassium may be obtained.



In this process six atoms of chlorine act on six atoms of potassium, thus,—

$$6Cl+3K_2O = 5KCl+KClO_3.$$

Any free chlorine is expelled from the solution by heat.

If carbonate of potassium be used instead of caustic potass, the carbonic acid is displaced with effervescence.

$$3K_2CO_3+6Cl = 5KCl+KClO_3+3CO_2.$$

14. Chlorate of potassium crystallizes in anhydrous four and six-sided scales, of a pearly lustre, soluble in twenty parts by weight of cold, and two of boiling water. It is much used for forming lucifer matches, which may be distinguished from simple phosphorus matches by the crackling noise they make when rubbed.

Ex. b.—See *Ex. f.*, page 64.

Ex. c.—Rub together in a dry mortar one grain of chlorate of potassium and half a grain of sulphur—a succession of explosions, like the cracking of a whip, will take place, accompanied with a purplish light.

Ex. d.—Mix *carefully* together two grains of chlorate of potassium and one of sulphur; wrap the mixture in a piece of strong paper, and strike it with a hammer on a stone or anvil; a loud report will be instantly produced.

Ex. e.—Mix together equal weights of chlorate of potassium and loaf sugar, both previously reduced to powder; place the mixture on a tile or plate, and let fall on it from a tube or rod, a single drop of strong sulphuric acid; a sudden and violent combustion will ensue.

Ex. f.—Put into a tall, narrow glass of water a little chlorate of potassium and a few shreds of phosphorus, and then, by means of a long dropping tube or funnel, pour to the bottom a few drops of strong sulphuric acid; violent action ensues, and the phosphorus burns under the water, the combustion being supported by the chlorine and gaseous compounds of chlorine and oxygen evolved by the decomposition of the salt.

N.B. Chlorate of potassium must never be kept ready mixed with inflammable bodies, for such mixtures sometimes explode spontaneously. In all experiments with these substances, only *very small quantities* should be used, and great caution must be exercised.

15. **Biborate of Sodium,** *borax,* $Na_2B_4O_7+10H_2O$.—This substance occurs in the waters of certain shallow lakes in Persia, India, and Thibet, and forms an incrustation on their banks. In its impure state it is imported into England under the name of *tincal.*

16. Borax crystallizes in six-sided prisms, which effloresce in dry air; it is soluble in about twenty times its weight of cold water. When heated, it swells up, owing to the escape of the water of crystallization, and becomes anhydrous; at a higher temperature it fuses, and on cooling forms a transparent glass. Its chief use is as a *flux,* and for soldering metals.

17. The object of employing borax in soldering, is to keep the metallic surfaces clean; this it does by dissolving the coating of oxide, and forming with it a glass which is fluid at a high temperature, and easily moved aside by the melted solder.

Borax is also much used as a flux in blowpipe analysis, owing to its forming colored glasses with various metallic oxides.

Ex. g.—Place a fragment of borax on a piece of charcoal, and heat it by the blow-pipe flame till it ceases to swell out, and has become

anhydrous; drop on it a little nitrate of cobalt, and then fuse it again, a *blue* glass will be the result. Oxide of silver similarly treated gives an opaque milk-white glass; oxide of manganese, a violet; oxide of nickel, an orange-red; &c.

Ex. h.—Dissolve some borax in hot water, soak some paper, or pieces of rag in the solution, and let them dry. It will be impossible to *inflame* them, though they will gradually burn away.

18. The **Silicates**, or compounds of silica with various bases, M_4SiO_4, form a large class of compounds, for the most part insoluble in water, and variously colored. They make up a very considerable portion of the earth's crust. The different kinds of glass and imitation jewels are artificial silicates.

19. Talc, French chalk, and soapstone, are silicates of magnesia, Mg_2SiO_4.

Serpentine is a double silicate of magnesia and iron, $MgFeSiO_4$.

Clay is Silicate of Alumina, Al_43SiO_4, combined with more or less oxide of iron.

The several varieties of *Feldspar* consist of silicate of alumina, combined with silicate of potass, soda, &c.

$$Al_43SiO_4 + K_4SiO_4; \; \textit{or}, \; AlKSiO_4.*$$

See also *Chap. XVIII.* § 8, and *Chap. XXII.* §§ 6, 7.

20. Silicate of potassium and sodium form a soluble glass (*Chap. XVII.* § 10).

Common window glass is a silicate of sodium and calcium.

Crown glass and *plate glass* are silicates of potassium and calcium.

Flint glass is a silicate of potassium and lead.

Bottle glass is a silicate of aluminum, iron, and potassium or sodium, the colour being due to the impurities of the materials.

Various metallic oxides are added to the melted silicates, to produce glass of different colours.

* If silicic acid be taken as dibasic (H_2SiO_3) these formula will be: M_2SiO_3; $MgSiO_3$; $MgFe2SiO_3$; Al_23SiO_3; $Al_23SiO_3 + K_2SiO_3$, *or* $AlK2SiO_3$.

21. **The Chromates,** $M_2C_2O_4$, present a great variety of beautiful colours, and are therefore much used as dyes and pigments, or as tests. They are, for the most part, insoluble in water.

22. **Chromate of Potassium,** K_2CrO_4, is a valuable salt, by means of which most of the other chromates may be prepared. It forms crystals of a beautiful yellow colour, soluble in about twice their weight of water at 60°. When strongly heated, it acquires a red colour, but becomes yellow again on cooling. When fused it loses oxygen, and becomes green; green oxide of chromium being formed, Cr_2O_3.

23. Chromate of potassium is prepared from *chrome-iron ore*, (a native compound of sesqui-oxide of chromium and protoxide of iron, Cr_2O_3+FeO), by strongly heating it either in a crucible, or a reverberatory furnace, with about one-fourth its weight of nitrate of potassium, and then dissolving out the soluble matter by means of water. The solution yields, by evaporation, anhydrous crystals of chromate of potassium, which are isomorphous with those of sulphate of potassium.

In the above process, the nitrate of potassium is decomposed, giving a portion of its oxygen to each of the metallic oxides, whereby the protoxide of iron, FeO, becomes the *per-* or *sesqui*-oxide, Fe_2O_3; while the sesquioxide of chromium, Cr_2O_3, becomes chromic acid, H_2CrO_4, which immediately combines with the potassium to form chromate of potassium, K_2CrO_4.

24. **Bichromate of Potass,** $K_2Cr_2O_7$, is obtained by removing half the potassium of the neutral chromate, by the action of sulphuric, nitric, or acetic acid: thus,—

$$2(K_2CrO_4)+H_2SO_4 = K_2SO_4+K_2Cr_2O_7+H_2O.$$

If sulphuric acid be used, the sulphate of potassium formed is not readily separated from the bichromate. The purest bichromate of potassium is obtained by using acetic acid.

25. Bichromate of Potassium is much employed in dyeing and calico printing. It forms beautiful red crystals, soluble in 10 parts of water at 60°. At a red heat it fuses without decomposition, but at a white heat, half the acid is decomposed, oxygen is evolved, and a mixture of sesquioxide of chromium and neutral chromate of potass remains:—

$$2(K_2Cr_2O_7) = Cr_2O_3 + 2(K_2CrO_4) + 3O.$$

26. **Chromate of Lead,** *Chrome yellow*, $PbCrO_4$.—This substance, which is much used by dyers and painters, is formed by adding a solution of chromate or bichromate of potass to one of acetate or nitrate of lead. It is insoluble in water, but soluble in nitric acid or in potass.

When chromate of lead is boiled with caustic potass or lime water, one-half the acid is withdrawn, and an orange-coloured subchromate of lead remains, $2PbCrO_5$ (*chrome orange*).

$$2(PbCrO_4) + K_2O = K_2CrO_4 + 2PbCrO_5.$$

When chromate of lead is gradually added to saltpetre in a state of fusion, at a gentle heat, a beautiful red subchromate is formed, termed *chrome red.*

Ex. i.—Dip some pieces of white linen or paper in acetate of lead and allow them to dry; then dip them in a solution of chromate of potassium; they will be dyed yellow; pass them into some boiling lime water, they will become orange-coloured.

Ex. k.—Prepare a series of glasses containing solutions of various metallic salts in distilled water; add to each a few drops of solution of chromate of potassium, and notice the colours of the precipitates formed, which will be chromates of the several metals.

27. **Arsenite of Potassium,** K_3AsO_3, is formed by boiling arsenious acid (common white arsenic) in water, with an equal weight of carbonate of potassium.

The alkaline arsenites are soluble and uncrystallizable.

28. **Arsenite of Copper,** $CuHAsO_3$. — This beautiful green powder, known as *Scheele's green*, is the precipitate formed by adding arsenite of potassium to sulphate of copper.

Ex. l.—Add a few drops of a solution of arsenious acid to a small quantity of a solution of the sulphate, nitrate, or acetate of copper. The liquid remains quite transparent and colourless, the arsenious acid not having so great an affinity for the oxide of copper as the acid with which it is already combined. If a small quantity of an alkaline solution (potass or its carbonate) be now added, the alkali will unite with the acid of the salt employed, and remain in solution, and the arsenious acid combining with the oxide of copper, will form arsenite of copper, which is insoluble in water, and is precipitated of a grass-green colour.

29. The other metallic arsenites may be formed by double decomposition, in a similar manner to arsenite of copper.

Arsenite of Silver, Ag_3AsO_3, is white, quickly changing to yellow, and finally to brown.

Ex. m.—See *Ex. k., Chap. XXV.*

30. **Arseniate of Potassium,** K_2HAsO_4, is formed by adding arsenic acid to carbonate of potass; it is uncrystallizable and deliquescent.

If the arsenic acid be in excess, a **Binarseniate of Potassium,** KH_2AsO_4, will be formed, which crystallizes on evaporation, and can thus be separated from the former.

Arsenic acid, heated in *earthenware* tubes, &c., soon corrodes them by its action on the alumina with which it combines, forming **Arseniate of Aluminum.**

31. Many of the arsenites and arseniates are found native in various metallic ores.

The precipitate of **Arseniate of Silver,** Ag_3AsO_4, formed by adding arseniate of potassium to nitrate of silver, is highly characteristic of arsen*ic* acid, being *red*, while the *arsenite of silver* is *yellow.*

Ex. n.—Add arseniate of potassium to a solution of the nitrate of silver. The arseniate of the metal will be thrown down as an insoluble powder of a *brick-red* colour.

32. The arseniates and arsenites may be recognised by the garlic-like odour they give off when heated in a tube with charcoal.

The *arsenites* when heated alone, generally lose part of

their acid, which sublimes; the *arseniates*, under the same circumstances, mostly retain their acid.

The *arsenites* when dissolved in dilute hydrochloric acid, give a yellow cloud of sulphuret of arsenic with sulphuretted hydrogen, but the *arseniates* do not.

EXERCISES ON CHAPTER XXXI.

1. How are the phosphates divided?
2. What is phosphate of sodium?
3. How may sub-phosphate of sodium be obtained?
4. How may acid-phosphate of sodium be obtained?
5. How may pyro-phosphate of sodium be obtained?
6. How may meta-phosphate of sodium be obtained?
7. What is microcosmic salt?
8. Illustrate the theory of replacement or substitution.
9. What is phosphate of calcium?
10. Why are burnt bones useful as manure?
11. Where is phosphate of lime found native? For what is it used?
12. State some general properties of the chlorates.
13. How is chlorate of potassium prepared?
14. What are the nature and properties of chlorate of potassium?
15. Whence is borax obtained?
16. What are the properties of borax?
17. For what is borax used?
18. Describe the silicates.
19. State the composition of talc, serpentine, clay, feldspar, and granite.
20. Name and give the composition of some artificial silicates.
21. Why are the chromates valuable?
22. Describe the chromate of potassium.
23. How is chromate of potassium prepared?
24. What are the characteristics and uses of bichromate of potassium?
25. How is bichromate of potassium prepared?
26. What are *chrome-yellow*, *orange-chrome*, and *chrome-red*, and how may they be formed?
27. What is arsenite of potassium?
28. What is *Scheele's green?*
29. How may the arseniates be formed?
30. Give the composition of arseniate and binarseniate of potassium.
31. State a characteristic test of *arsenic acid*.
32. How may the arsenites and arseniates be recognised and distinguished?

CHAPTER XXXII.

HALOID SALTS.

CHLORIDES.

1. The Chlorides may be formed, either by the direct union of their elements (*a*), or by the action of hydrochloric acid on a metal (*b*), or metallic oxide (*c*), the hydrogen of the acid being replaced by its equivalent of metal.

(*a*). As in the combustion of antimony, *Chap.* xvi. *Ex. g.* $Sb + 3Cl = SbCl_3$.

(*b*). As when zinc is added to strong hydrochloric acid, and hydrogen gas is given off. $Zn + 2HCl = ZnCl_2 + 2H$.

(*c*). As when common salt is formed by adding hydrochloric acid to soda (*double decomposition*). $Na_2O + 2HCl = 2NaCl + H_2O$.

2. The attraction of chlorine for metals is greater than that of oxygen; hence when a metallic oxide is heated in chlorine, oxygen gas is evolved, and a chloride formed.

All the metals unite with chlorine, some at common temperatures, others when more or less heated; some very slowly, others with great energy and rapidity. Minute division accelerates the combination.

Ex. a.—Perform *Ex. g.*, *Chap.* xvi.

3. Most of the metallic chlorides are soluble in water, and may be readily crystallized; some are wholly insoluble, while others exist only in the liquid state.

The insoluble chlorides may be obtained by adding a soluble chloride to the soluble metallic salts.

Ex. b.—Add a drop of nitrate of silver to a clear solution of common salt, a milkiness will instantly appear, owing to the formation of the insoluble chloride of silver.

$$AgNO_3 + NaCl = NaNO_3 + AgCl.$$

Hence nitrate of silver is an excellent test for hydrochloric acid.

4. The chlorides may, for the most part, be decomposed by hydrogen at a red heat, hydrochloric acid being formed, and the metal *reduced*. The *noble metals* may be reduced from their chlorides by heat alone.

The chlorides may likewise be decomposed by the action of a strong acid, as the sulphuric, (see the mode of obtaining hydrochloric acid, *Chap*. xvi. § 14,) or by a metal which has a greater affinity for chlorine than the one with which that element is already in combination: thus chloride of silver is decomposed by zinc, &c.

5. **Chloride of Sodium,** *common salt,* NaCl.—This substance is found, partly in the solid form, as *rock-salt,* and partly dissolved in the waters of *saline springs,* but most abundantly in *sea water.*

Sea water contains on an average about 2½ per cent. of salt. Its nauseous, bitter taste arises from the salts of magnesium which exist in it.

6. Immense quantities of salt are obtained by evaporating the waters of the ocean, or of brine springs, either by exposure to the sun and air, or more generally, by artificial heat.

The salt separates in crystals, while the impurities remain in the small portion of liquid which is not evaporated, (*mother-water*).

7. Rock-salt is found in beds of great thickness, in Cheshire, Gloucestershire, Staffordshire, Spain, Poland, and many other parts of the world. To purify it for general use, it is dissolved in water, and the solution evaporated and crystallized, as already described.

8. Chloride of sodium is colorless, inodorous, transparent, and brittle; its specific gravity is about 2.125; it crystallizes in anhydrous cubes, which are deliquescent in moist air.* Contrary to the general rule, salt is just as

* When *perfectly pure,* chloride of sodium is unaltered by exposure to the air; but it is generally more or less deliquescent, from containing chlorides of magnesium and calcium.

soluble in cold as in boiling water; in fact, water at 32° dissolves rather more than water at 60°. It is insoluble in alcohol. Chloride of sodium fuses at a red heat, without decomposition, and at a still higher temperature it is sublimed.

Ex. e.—Mix together equal parts of snow and salt. The two will unite and form a liquid—colder than either of the two before mixing. A still greater degree of cold would be produced by using chloride of potassium instead of salt.

9. Salt exists in small quantities in many plants, and is hence sometimes useful as a manure. It also exists in animal bodies and in the blood, and forms an indispensable constituent of the food of man and of most animals.

Salt is used for packing and preserving meat, as it retards or prevents putrefaction. It is also employed in the preparation of chlorine, hydrochloric acid, carbonate of sodium, chloride of ammonium, and several other salts.

10. **Chloride of Potassium**, KCl, was formerly used in medicine: it is one of the constituents of sea water, and may be formed artificially, either by heating potassium in chlorine or in gaseous hydrochloric acid, or by dissolving potass (or its carbonate) in that acid.

Chloride of potassium is the substance left in the retort when oxygen is prepared from chlorate of potassium, *Chap.* iv. § 6. It is also obtained in considerable quantities as a residue in several chemical manufactures, especially in that of the salt last mentioned. It is sometimes used in alum-making.

11. When chlorine gas is passed through a solution of potass (or its carbonate), a part of the potass is decomposed; its potassium combines with chlorine, forming chloride of potassium, while its oxygen combines with another portion of the chlorine, forming chloric acid; this unites with the undecomposed potass, and forms chlorate of potassium.

The chlorate of potassium may be readily separated by

crystallization, while the chloride of potassium which remains in the mother liquor may be obtained pure by evaporation to dryness and exposure to a dull red heat.

$$6Cl+6KHO = KClO_3+3H_2O+5KCl.$$

Ex. d.—Add 1 oz. of the powdered chloride of potassium to 4 oz. of water; stir them together, and a great degree of cold will be produced, equal to about 8 times as great a reduction as with common salt.

12. **Chloride of Ammonium,** *sal-ammoniac,* NH_4Cl, was formerly obtained from Egypt, where it was prepared by the distillation or rather sublimation of camels' dung. Pliny says the ancients called it *sal-ammoniacum*, because it was found near the temple of Jupiter Ammon, in Libya. It often occurs native in the neigbourhood of volcanoes, as in the Solfatara, near Naples, among the lava of Etna, and in the country of Bokhara.

13. Chloride of ammonium is now prepared on a large scale, by saturating with hydrochloric acid the ammoniacal liquid produced in the manufacture of coal gas, or during the distillation of bones and other refuse, in the preparation of animal charcoal. The acid neutralizes the free alkali and decomposes the carbonate and sulphide of ammonium, with evolution of carbonic acid and sulphuretted hydrogen in the form of gas. The solution is then evaporated to dryness, and the salt purified by sublimation.

$$(NH_3),H_2O+HCl = (NH_4)Cl+H_2O.$$
$$(NH_4)HS+HCl = (NH_4)Cl+H_2S.$$
$$(NH_4)HCO_3+HCl = (NH_4)Cl+H_2O+CO_2.$$

14. Chloride of ammonium is likewise largely prepared by decomposing sulphate of ammonium by common salt, thus—

$$(NH_4)_2SO_4+2NaCl = Na_2SO_4+2(NH_4)Cl.$$

The sulphate of ammonia used in this process is obtained by adding sulphuric acid to the ammoniacal liquid mentioned in § 13, and purifying the product.

15. **Sal-ammoniac** is a white salt, having a sharp saline taste; it is soluble in water, and crystallizes in anhydrous cubes or octohedra; it may be sublimed without decomposition, and when sublimed is tough, fibrous, and difficult to be powdered.

It is used in various processes of metallurgy,* as, for instance, in tinning, to prevent the oxidation of the surface of the copper; it is also employed in dyeing and medicine, as well as in many other arts.

Aqua regia, which is used as a solvent for gold, is often formed by adding nitric acid to this salt, instead of by directly mixing the nitric and hydrochloric acids.

Ex. e.—See *Ex. d.*, *Chap.* xxiv.

16. **Chloride of Calcium**, $CaCl_2$, is found in sea-water, and in some saline springs and mineral waters. It is formed by heating lime in chlorine, or by dissolving carbonate of calcium in hydrochloric acid. It is also a by-product in several chemical manufactures, as for instance in that of carbonate of ammonium, which is made by subliming a mixture of sal-ammoniac and chalk; *Chap.* xxviii., § 14.

17. Chloride of calcium has a strong attraction for water, being one of the most deliquescent salts known. Its taste is bitter and disagreeable. It is soluble in alcohol, and much heat is evolved during the solution. When fused it has phosphorescent properties, and has been termed Homberg's phosphorus, since he first observed them. Owing to its affinity for water, fused chloride of calcium is largely used for drying gases.

Ex. f.—Put half an ounce of chloride of calcium, with the same quantity of acetate of lead, or nitrate of sodium, into a mortar—rub them together. These salts will part with their water of crystallization, and become fluid.

18. There is a substance commonly sold under the name of "chloride of lime," which is valuable as a

* *Metallurgy*=the art of working metals.

bleaching powder and a *disinfectant.* The name, "chloride of lime," is incorrect, as chlorine, a simple body, cannot combine with lime, which is a compound one. It may be more correctly regarded as a hypochlorite of calcium, or rather as a mixture of chloride of calcium, hypochlorite of calcium, and lime still uncombined. It is a white powder, somewhat soluble in water, and having an odour similar to that of chlorine.

19. **Hypochlorite of Calcium** is formed by passing chlorine gas through a mixture of lime and water (*lime white*). The chlorine combines with both constituents of the lime, forming with the metal, chloride of calcium; and with the oxygen, hypochlorous acid. This acid combines with another portion of the lime, and forms hypochlorite of calcium.

$$2CaH_2O_2 + 4Cl = CaCl_2 + Ca2ClO + 2H_2O.$$

20. The solution of so-called chloride of lime has in itself very slight bleaching properties, and these principally through the action of the carbonic acid of the air, which displaces the chlorine. The bleaching effect is instantly produced by a drop or two of one of the stronger acids, owing to their immediate combination with the lime and the liberation of chlorine. (See *Chap.* xvi. §§ 11, 12.)

Ex. g.—Put a few drops of a solution of indigo into a clear solution of the chloride of lime, and add a very little hydrochloric or nitric acid; the color of the indigo will be immediately destroyed. The same may be tried with the solution of any other animal or vegetable coloring matter

Ex. h.—See *Ex. b.*, page 113.

Ex. i.—Dip a strip of red calico alternately in a solution of hypochlorite of lime, and in some slightly acid water. Chlorine will be set free in the fibres of the cloth, and the color will be very speedily discharged.

Stains from port wine, fruit juice, ink, &c. may be readily removed from linen or cotton in this way, without injuring the material, if dilute solutions only be used, and sufficient time be allowed.

21. **Chloride of Barium,** $BaCl_2$, is a greyish white salt, soluble in water, but insoluble in alcohol. It is much used

as a test for sulphuric acid, with which it forms an insoluble white precipitate of sulphate of barium. It may be obtained in an anhydrous state, by dissolving native carbonate of barium in hydrochloric acid, evaporating the solution to dryness, and fusing the residue in a covered platinum crucible. If the solution be only partially evaporated, and left to cool, colorless, transparent crystals will be deposited, containing two equivalents of water, $BaCl_2+2H_2O$.

22. **Chloride of Magnesium,** $MgCl_2$, is the principal ingredient in the liquid which remains after the crystallization of common salt from sea-water. It is extremely deliquescent, very soluble in water and alcohol, and has a nauseous, bitter taste. It cannot be obtained by evaporating its aqueous solution to dryness, for the last portions of water are decomposed by the concurrent attraction of chlorine for hydrogen and of magnesium for oxygen, so that hydrochloric acid is given off and magnesia alone remains.

$$MgCl_2+H_2O = MgO+2HCl.$$

If, however, equal parts of hydrochloric acid be neutralized, one with magnesia and the other with ammonia, and the hydrated chloride of magnesium and chloride of ammonium so formed be evaporated to dryness, and raised to a red heat, the water is driven off, the chloride of ammonium sublimes, and chloride of magnesium remains in a fused state.

23. **Chloride of Zinc,** $ZnCl_2$, is readily prepared by dissolving the metal in hydrochloric acid. It is deliquescent and soluble both in water and alcohol. Its aqueous solution, known as *Sir W. Burnet's disinfecting solution*, is much used as an *antiseptic* and a *disinfectant*, since it arrests the decay of organic matter, and decomposes sulphuretted hydrogen and sulphide of ammonium.

24. **Protochloride, or Dichloride of Tin,** $SnCl_2$, called in commerce, *tin-salts*, is a grey, semi-transparent, crystalline substance. It is used as a test for mercury and gold, these metals being precipitated by it in a state of minute division.

Chloride, or Perchloride of Tin, $SnCl_4$, the *fuming liquor of Libavius*, is a transparent, colourless liquid, prepared by heating bichloride of mercury and powdered tin in a retort, and condensing the vapours that arise. A solution of this substance is much used in dyeing as a *mordant*, for fixing colours on cloth, &c.

Ex. k.—Place a slip of clean zinc in a solution of perchloride of tin, decomposition ensues, the chlorine unites with the zinc, forming chloride of zinc, while metallic tin is precipitated.

25. **Chloride of Copper, $CuCl_2$**, is formed by dissolving the protoxide of copper in hydrochloric acid. It forms green crystals, soluble in water and alcohol ; the alcoholic solution burning with a green flame. When heated, it parts with half its chlorine, and becomes—

Sub-chloride of Copper, $CuCl$, a white substance, but slightly soluble in water, and prone to oxidize in moist air, becoming changed into a green powder (*Brunswick green*), very much used by house-painters for common out-door work.

No green color is so cheap as this, and it is the one most commonly employed. It is a good color when first laid on, but after a time, it becomes partly changed into oxide of copper, and assumes a blackish-green tint, especially if it be in the winter season.

Ex. l.—Write on a sheet of paper with a solution of chloride of copper (using a quill pen), nothing will be seen while the paper is damp, but if heated, the writing will appear yellow.

Chloride of Cobalt, $COCl_2$, used in the same way, will produce a fine blue, though pale pink while moist ; similarly,

Acetate of cobalt yields a fine green.

26. **Subchloride, or Dichloride of Mercury**, *Calomel*, $HgCl$. This valuable salt may be obtained in several ways—

1. By heating the metal in chlorine gas, the *insoluble* portion of the residue is the sub-chloride.
2. By adding an atom of chloride of sodium to an atom of the *sub-* or *proto*-nitrate of mercury.

$$HgNO_3 + NaCl = NaNO_3 + HgCl.$$

3. By dissolving *proto*xide or *grey* oxide of mercury, Hg_2O, in hydrochloric acid.

$$Hg_2O + 2HCl = H_2O + 2HgCl.$$

4. By triturating together, mercury, *per*sulphate of mercury, and common salt, and subliming the mixture. Subchloride of mercury is vaporized and condensed, while sulphate of sodium remains.

$$Hg + HgSO_4 + 2NaCl = Na_2SO_4 + 2HgCl.$$

The *per*sulphate of mercury is obtained by boiling metallic mercury in sulphuric acid and evaporating to dryness.

If *proto*sulphate of mercury, Hg_2SO_4, could be obtained as readily as the *per*sulphate, it would be employed instead of it, and the metallic mercury might be dispensed with, in the above process. The mechanical mixture, however, of the *per*sulphate and mercury, is equivalent, in result, to the actual *proto*sulphate.

27. **Perchloride**, or more correctly, **Chloride of Mercury**, *Corrosive sublimate*, $HgCl_2$, is also to be obtained in several ways.

1. The *soluble* portion of the mass formed by heating mercury in chlorine is the perchloride.
2. When *per*oxide or *red* oxide of mercury, HgO, is dissolved in hydrochloric acid, the results are water and perchloride of mercury, which may be crystallized, on evaporation.

$$HgO + 2HCl = H_2O + HgCl_2.$$

3. By subliming a mixture of *per*sulphate of mercury and common salt.

$$HgSO_4 + 2NaCl = Na_2SO_4 + HgCl_2.$$

Refer to *Chap.* xxv. §§ 6, 7, 8.

28. Pure *calomel* is a white, heavy, *insoluble*, tasteless powder, which is volatilized by heat, subliming into a yellowish white, crystalline mass. Potass, soda, and lime-water decompose it, grey suboxide of mercury, Hg_2O, being separated. Calomel is much used in medicine as a purgative, but must be administered with care.

29. *Corrosive sublimate* is a white, semitransparent, crystalline substance, *soluble* in water, of a nauseous, acrid taste, and highly poisonous. Potass, soda, or lime-water, added to a solution of corrosive sublimate, throws down a precipitate of finely-divided oxide of mercury, HgO.

Corrosive sublimate is a powerful *antiseptic*, and is much used for preserving timber, &c. from decay (*Kyanizing process*).

30. **Chloride of Silver**, $AgCl$, may be formed by adding any soluble chloride, as common salt, to any soluble salt of silver, except the hyposulphite, when a white, curdy precipitate falls, which is quite insoluble in water, but readily dissolved by ammonia, cyanide of potassium, or hyposulphite of soda, &c.

Chloride of silver is decomposed, and becomes black on exposure to light, very slowly if pure, but more quickly if organic substances be present. When carefully heated, it melts, and forms, on cooling, a grey, horn-like mass. In this state, it is frequently found native, and is termed *horn-silver*.

EXERCISES ON CHAPTER XXXII.

1. How are the chlorides formed?
2. How are the metals affected by chlorine?
3. How are the metallic chlorides acted on by water? Illustrate the formation of an insoluble chloride.
4. How may the chlorides be decomposed?
5. Where is chloride of sodium found?
6. How is common salt obtained?
7. Whence is rock-salt obtained, and how is it purified
8. State the properties of chloride of sodium.
9. Name some of the uses of common salt.
10. Whence is chloride of potassium obtained?
11. Describe the reaction of chlorine on a solution of potass.
12. What is *sal-ammoniac?* Whence does it derive its name? Where is it found native?
13. How is chloride of ammonium generally prepared?

14. Shew the reaction of common salt on sulphate of ammonium?
15. What are the chief characteristics and uses of chloride of ammonium?
16. What is chloride of calcium?
17. State the nature and uses of chloride of calcium.
18. What is common bleaching powder?
19. How is hypochlorite of lime prepared?
20. Whence does hypochlorite of lime derive its bleaching properties?
21. State the composition and uses of chloride of barium.
22. What is chloride of magnesium?
23. What are the uses of chloride of zinc?
24. Describe the chlorides of tin.
25. Describe the chlorides of copper.
26. How may subchloride of mercury be formed?
27. How may corrosive sublimate be prepared?
28. State the properties and uses of calomel.
29. What are the uses and properties of corrosive sublimate?
30. Describe the nature and properties of chloride of silver.

CHAPTER XXXIII.

HALOID SALTS—*continued.*

IODIDES, BROMIDES, SULPHIDES, ETC.

1. **The Iodides and Bromides** have many properties in common with the chlorides, and may, for the most, be obtained in a similar manner—either by the direct action of the combining elements, as when a metal is burned in the vapor of iodine or bromine, — or by the action of hydriodic or hydrobromic acid, on a metal or metallic salt; the hydrogen of the acid being replaced by its equivalent of metal.

Ex. a.—See *Chap* xx., *Ex. f.*

2. **Iodide of Potassium, KI.**—When iodine is added to a strong solution of caustic potass, a mixture of iodate and iodide of potassium is formed as in the analogous case of chlorine, (*a*) (*Chap.* xxxii. § 11); on evaporating the solution to dryness, and cautiously raising it to a red heat, the iodate of potassium parts with its oxygen, and is entirely charged into iodide of potassium (*b*).

$$(a)\ 6I + 6KHO = KIO_3 + 5KI + 3H_2O.$$

$$(b)\ KIO_3 = KI + 3O.$$

3. If, instead of fusing the product of the evaporated solution (*a*), alcohol be poured upon it, the iodide of potassium will be dissolved, and a salt insoluble in alcohol will remain, which is the **Iodate of Potassium,** KIO_3, analogous to chlorate of potassium.

Ex. b.—Perform *cautiously Ex.* c. and *d.*, and *Chap.* xxx., using the *iodate* instead of the *chlorate* of potassium.

4. Another, and perhaps better mode of obtaining iodide of potassium, is by decomposing the iodide of zinc (or iron) by carbonate of potassium.

The iodide of zinc (or iron) is made by gently heating together iodine, water, and scraps of zinc (or iron). The resulting iodide is filtered, and exactly decomposed by carbonate of potassium, carbonate of zinc (or iron) is precipitated, and iodide of potassium remains in solution. The latter is separated by filtration, and then evaporated, and crystallized.

$$ZnI_2 + K_2CO_3 = ZnCO_3 + 2KI.$$

5. Iodide of potassium is important as a medicinal compound (see page 118). It crystallizes in anhydrous non-deliquescent cubes, which fuse readily when heated, and are very soluble in water. Most of the metallic iodides

may be obtained by adding iodide of potassium to the soluble salts of the metals.

Ex. c.—Put a little starch paste into water, add a few drops of iodide of potassium, no effect will be produced, for the iodine being already in combination, cannot unite with the starch. Add a little dilute acid or chlorine water, the iodine will be at once set free, and combining with the starch, form its characteristic *blue* iodide. *Chap.* xvi. § 22, *Ex. o.*)

Ex. d.—Fill a series of test glasses with the following solutions in distilled water, add to each a few drops of iodide of potassium, and note the colors of the iodides produced.

Sulphate of Copper	gives	Iodide of Copper		*brown.*
Acetate of Lead	,,	,,	,, Lead	*yellow.*
Nitrate of Bismuth	,,	,,	,, Bismuth	*orange.*
,, ,, Silver	,,	,,	,, Silver	*primrose.*
Sub-nitrate of Mercury	,,	Sub-iodide	,, Mercury	*yellow.*
Perchloride of ,,	,,	Per-iodide	,, ,,	*scarlet.*

6. **Sub-Iodide, or Di-iodide of Mercury,** HgI, is formed by adding iodide of potassium to the subnitrate of mercury; it appears as a dirty, greenish yellow, insoluble precipitate.

Per-Iodide, or, more correctly, **Iodide of Mercury,** HgI_2, is formed by adding iodide of potassium to a solution of corrosive sublimate, a yellow precipitate falls, changing immediately to a most beautiful scarlet. On a further addition of iodide of potassium, the precipitate is dissolved, and the liquid becomes clear. The scarlet precipitate, when dried, has been used as a water colour (*geranium red*), but, unfortunately, it is not permanent.

7. The per-iodide of mercury affords another illustration of *dimorphism*, (*Chap.* xxvii. § 24), attended with change of color; for when the scarlet precipitate is heated it sublimes, yielding a quantity of minute, brilliant, yellow crystals. On touching this yellow incrustation with a hard substance, as the point of a needle, it instantly becomes red at the point of contact, and the color gradually spreads through

the whole mass, till every particle has regained its original scarlet.

8. **Teriodide of Phosphorus**, PI_3. See *Chap.* iii. *Ex. d.*

9. **Teriodide of Nitrogen**, NI_3. See *Chap.* xvi. § 24.

10. **Bromide of Potassium**, KBr, is a white, fusible, and very soluble salt, similar in appearance and general properties to iodide of potassium, and prepared in a similar manner. It is only used in photography.

Ex. e.—Perform *Ex. c.*, using the bromide instead of the iodide of potassium, a *pale orange* bromide of starch will be produced.

11. **Fluoride of Calcium**, *fluor-spar*, CaF_2.—The combinations of fluorine with the other elements are very few, and those few but little understood. The most important is the beautiful mineral termed *fluor-spar*, which is found in many parts of Derbyshire, crystallized in regular cubes or octohedra, of a white, green, rose, or purple color. It fuses at a red heat, and is sometimes used as a flux for metallic ores, hence its name "*fluor* spar."

Hydrofluoric acid is obtained from it by the action of sulphuric acid, as shewn, *Chap.* xvi. § 27.

$$CaF_2 + H_2SO_4 = CaSO_4 + 2HF.$$

12. The metallic *sulphurets* or *sulphides*, MS, may be prepared by rubbing or melting together sulphur and a metal, *Chap.* iii. *Ex. e*; *Chap.* xx. *Ex. e*;—by adding sulphuretted hydrogen, or sulphide of ammonium, to a metallic oxide or salt, *Chap.* xiv. *Ex. o, p, q*;—by heating metallic sulphates with charcoal or carbonaceous matter, to remove the oxygen of the salt, *Chap.* xxix. § 6.

13. Some of the metallic sulphurets undergo no change when heated in close vessels, as the sulphurets of potassium and sodium; others sublime unaltered, as the sulphurets of arsenic and mercury, *Chap.* xxiv. § 18; *Chap.* xxv. § 9;—others lose a portion of their sulphur, and if air have access, give off sulphurous acid, the metal passing

into the state of oxide (*roasting*), as the sulphurets of copper and lead, *Chap.* xxiii. §§ 4, 15;—others are completely reduced to the metallic state, as the sulphurets of gold and platinum.

14. The metallic sulphurets are, for the most part, opaque, brittle solids. They are all insoluble in water, except those of the metallic bases of the alkalies and alkaline earths.

15. Several of the metallic sulphurets occur abundantly in nature; the most common are, the sulphurets of iron, lead, antimony, copper, zinc, arsenic, mercury, and silver. These, with the exception of the first, have been noticed in the accounts of the several metals.

16. The sulphuret of zinc is white; that of arsenic, yellow; that of antimony, orange. The rest of the insoluble sulphurets are black.

17. **Protosulphuret of Potassium,** K_2S, is formed by heating potassium with sulphur, *Chap.* xx. *Ex. e.*, or by heating sulphate of potassium with charcoal. There are three other sulphurets of potassium, viz., the **Bisulphuret,** KS; the **Tersulphuret,** K_2S_3; and the **Pentasulphuret,** K_2S_5. An impure variety of the latter was formerly called *liver of sulphur*, from its color. If an acid be added to a solution of it, sulphur will be precipitated in a state of the finest division, giving the liquid the appearance of milk (*milk of sulphur*).

18. **Protosulphuret of Sodium,** Na_2S, is formed in a similar manner to the preceding; see also *Chap.* xxvii. § 10. It is interesting, as forming one of the ingredients in the mineral termed *lapis-lazuli*, from which *ultramarine* is prepared.

"Artificial ultramarine is obtained by heating sulphuret of sodium with silica, aluminum, and iron."

There are several other unimportant sulphurets of sodium.

19. Several sulphurets of ammonium may be obtained by distilling the sulphurets of potassium or sodium with chloride of ammonium.

When two volumes of ammoniacal gas are mixed with one volume of sulphuretted hydrogen, and condensed by cold, a neutral **Sulphuret of Ammonium,** NH_4S, is obtained in extremely volatile deliquescent crystals.

20. **Sulphide of Ammonium,** or as it is frequently called, **Hydrosulphate,** or **Hydrosulphuret of Ammonia,** $(NH_4)HS$, is a *double sulphuret of ammonium and hydrogen,* formed by passing a current of sulphuretted hydrogen gas through liquid ammonia, till the latter is saturated with the gas, which may be known by the solution ceasing to give a precipitate with sulphate of magnesium. The sulphide of ammonium so prepared is at first nearly colorless, but as it gradually absorbs oxygen from the air, it becomes yellow after a time, from the presence of excess of sulphur, which may be precipitated by the addition of an acid.

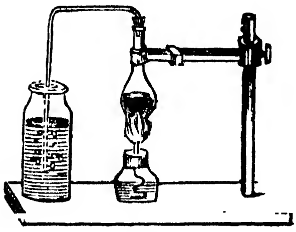


Fig. 77.

The sulphide of ammonium is a useful test, as it precipitates many metals from their solutions.

21. There are several of the double sulphurets which have been termed *sulphur-salts,* for one of the sulphurets seems to act the part of an acid (*sulphur-acid*), as sulphuretted hydrogen, sulphuret of carbon, of arsenic, of antimony, &c., while the other acts the part of a base (*sulphur-base*), as the proto-sulphuret of ammonium, and the proto-sulphurets of the metals of the alkalies and alkaline earths.

The sulphuret or sulphide of ammonium might thus be called a sulphur-salt, the sulphur-base, *neutral sulphuret of ammonium*, being combined with the sulphur-acid, ***sulphuretted hydrogen.***

22. **Proto-sulphuret of Iron,** *Magnetic pyrites*, FeS, is a blackish, brittle subtance, attracted feebly by the magnet, formed by heating iron filings and sulphur together in a crucible. It is found native in Cornwall, Wales, Norway, the Harz mountains, and in North America.

Ex. f.—Heat a rod of iron to whiteness, then rub it with a stick of roll sulphur, avoiding the fumes, and allowing the drops to fall into water. The black substance found at the bottom will be the proto-sulphuret of iron.

23. **Bisulphuret of Iron,** *Iron pyrites*, FeS_2, is a substance of a brass-yellow color, which occurs abundantly in every part of the world, and in different forms, derived from the cube. It is often found in nodules, which when broken have a radiated appearance; they are sometimes, though wrongly, supposed to be of meteoric origin, and have therefore been called *thunderbolts*.

The principal use of iron-pyrites is in the manufacture of sulphate of iron, *Chap.* XXIX. § 22.

24. **Bisulphuret of Carbon,** CS_2, is a transparent, colorless liquid, with a most disagreeable odor; it is exceedingly volatile, and produces by its evaporation an intense degree of cold. It boils at 110°F.

It burns with a blue flame, the products of combustion being carbonic acid and sulphurous acid gas. It freely dissolves sulphur and phosphorus, and the solution often deposits these substances in beautiful crystals, by spontaneous evaporation.

Ex. g.—Half fill a phial with water or with sulphuric acid, and surround it with a rag moistened with bisulphuret of carbon. The liquid will speedily be frozen.

Ex. h.—Dip a piece of paper in a solution of phosphorus in bisulphuret of carbon, and allow the moisture to evaporate, the phosphorus, being left in a state of minute division, will immediately inflame.

25. **Phosphuret of Calcium,** CaP_2, is a brown compound formed by passing the vapor of phosphorus over pieces of lime heated to redness in a porcelain tube.

Fig. 78.

Ex. i. — When put into water, this substance, if fresh made, gives off spontaneously-inflammable phosphuretted hydrogen. See *Chap.* xv. § 14, and *Ex. m.*

EXERCISES ON CHAPTER XXXII.

1. What are the general properties of the iodides and bromides, and how may they be obtained?
2. How is iodide of potassium prepared from iodine and potass?
3. How is iodate of potassium obtained?
4. How are the iodides of zinc and iron prepared, and how are they acted on by carbonate of potassium?
5. What are the properties and uses of iodide of potassium?
6. Describe the iodides of mercury.
7. What is *dimorphism,* and how does per-iodide of mercury illustrate it?
8. How is ter-iodide of phosphorus prepared?
9. Describe the preparation and properties of ter-iodide of azote.
10. What is bromide of potassium?
11. Describe the fluoride of calcium.
12. How may the metallic sulphurets be prepared?
13. How are the metallic sulphurets affected by heat?
14. Which are the soluble, and which the insoluble sulphurets
15. Name the chief native sulphurets.
16. State the characteristic colors of some of the sulphurets.
17. Describe the sulphurets of potassium.
18. Of what use is sulphuret of sodium?
19. How are the sulphurets of ammonium obtained?
20. Describe the preparation, properties, and uses of the common sulphide of ammonium (hydrosulphuret of ammonia)
21. Explain the constitution of sulphur-salts.
22. What is the proto-sulphuret of iron?
23. What is iron-pyrites, and for what is it used?
24. What are the properties of bisulphuret of carbon?
25. What is phosphuret of calcium?

CHAPTER XXXIV.

ORGANIC COMPOUNDS.

1. ORGANIC chemical compounds are compounds directly or indirectly derivable from organized bodies. They cannot, except in very rare instances, be formed by bringing their elements together, "but must either be derived *ready-made* from plants and animals, or must be prepared from such *ready-made* substances, which form, so to speak, the raw material."

Thus, sugar may be shewn to consist of carbon and water, *Chap.* iii. *Ex. n.*; yet we cannot unite carbon, oxygen, and hydrogen, so as to form sugar. Again, alcohol and vinegar contain the same three elements, but can only be obtained from sugar, as the raw material, by fermentation or chemical action.

2. The prevailing elements in organic compounds are carbon, oxygen, hydrogen, nitrogen, sulphur, and phosphorus.

Some organic bodies consist of only two elements,—carbon and oxygen; a great many are found to contain three,—carbon, oxygen, and hydrogen; not a few contain four,—carbon, hydrogen, oxygen, and nitrogen; very few contain five or six, though almost every element *may* occur in some organic compound or other.

3. Organic compounds consisting only of carbon, oxygen, and hydrogen (*non-nitrogenous compounds*), are easily distinguished from those which in addition, contain nitrogen, or nitrogen with sulphur or phosphorus (*nitrogenous compounds*), by burning them. The latter class will evolve an offensive odor like burnt feathers, or if heated with lime, will give off ammonia.

4. Organic compounds may be classed as *acids*, *bases*, and *indifferent* or *neutral* substances. The organic acids are chiefly contained in fruits, or in the sap of plants.

They are usually prepared by saturating the juices containing them with lime, evaporating the solution of the *lime-salt* to dryness, and afterwards decomposing it by sulphuric acid, which removes the base and liberates the organic acid, which can be separated either by filtering or by evaporation.

5. **Acetic Acid**, $C_2H_4O_2$, is readily formed by the fermentation of many animal and vegetable juices, when exposed to the action of the air at a moderate temperature. It is also a product of the distillation of vegetable matter, especially wood, in close vessels (*wood vinegar* or *pyroligneous acid.*)

Ordinary vinegar is dilute acetic acid; it cannot be concentrated by evaporation, for the acid is volatile as well as the water.

"The strongest acetic acid is obtained by distilling anhydrous acetate of sodium with three times its weight of concentrated sulphuric acid."

$$2Na(C_2H_3O_2) + H_2SO_4 = Na_2SO_4 + 2(H,C_2H_3O_2).$$

Acetic acid is monobasic and contains one replaceable atom of hydrogen. Its empirical formula is as above, $C_2H_4O_2$. Its rational formula may be written, $H,C_2H_3O_2$, or $\left.\begin{matrix} C_2H_3O \\ H \end{matrix}\right\} O$.

6. **Acetate of Lead**, *sugar of lead*, $Pb(C_2H_3O_2)_2$, is prepared by dissolving *litharge*, protoxide of lead, in strong vinegar, or else by exposing plates of lead to the vapor of vinegar, and dissolving the carbonate of lead so formed, (*Chap.* xxvii. § 27) in acetic acid. The acetate of lead of commerce is a white, crystalline substance, of a sweetish taste, somewhat resembling loaf sugar, whence the name *sugar of lead*. It is much used in calico-printing and dyeing; also by the painter, on account of its promoting the drying of oil colors by absorbing the greasy particles of the oil.

It is a violent poison, but is used medicinally, as an *external* lotion for the eyes, &c. (*Goulard's water.*)

7. **Acetate of Copper,** *crystallized verdigris,* $Cu(C_2H_3O_2)_2$.—When sheets of copper are exposed to the action of acetic acid a subacetate of copper, $Cu(C_2H_3O_2)$, is formed, known in commerce as *common verdigris;* if this be dissolved in hot acetic acid and allowed to crystallize slowly, it will deposit beautiful dark-green crystals of neutral acetate, used by painters under the name of *distilled* or *crystallized verdigris.*

8. **Oxalic acid,** $H_2, C_2O_4+2H_2O$, is found in the stems of rhubarb, in the sap of common and wood sorrel, and in some fruits, combined with potass or lime. It is prepared artificially in large quantities by the action of nitric acid on sugar, starch, and dextrine (*Chap.* iii. *Ex. q.*). The crystallized acid has very much the appearance of Epsom salts; it is very soluble, the solution being intensely sour, and highly poisonous; the proper antidote is chalk or magnesia.

When heated with sulphuric acid, oxalic acid is decomposed into carbonic acid and carbonic oxide, see *Chap.* xi. §§ 19. 20.

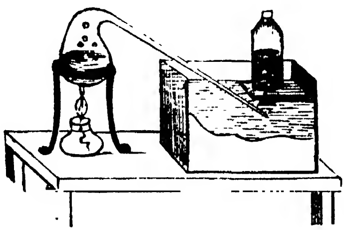


Fig. 79

9. **Oxalate of Ammonium,** $(NH_4)_2C_2O_4$, is prepared by neutralizing a hot solution of oxalic acid by carbonate of ammonia. It is used in chemical analysis to precipitate lime from its solutions, which it does in the form of a white powder, **Oxalate of Calcium,** Ca,C_2O_4.

10. Oxalic acid forms with peroxide of iron a very

soluble **Per-oxalate of Iron,** $Fe_2 3(C_2O_4)$, so that either the acid itself, or the **Binoxalate of Potassium** (*salt of sorrel*), KHC_2O_4, is used for removing *ink-stains* (tanno-gallate of iron), or *iron-mould* (per-oxide of iron), from linen, &c., for a soluble oxalate of iron is formed which can be easily washed away.

Ex. a.—Oxalic acid is the only acid which will decompose sulphate of lime. Add some solution of oxalic acid, or oxalate of ammonium, to lime-water, or any salt of calcium, oxalate of calcium will be precipitated as a white powder.

11. **Tartaric Acid,** $H_2,C_4H_4O_6$, is a *bibasic* acid, found in a free state in the juice of unripe grapes, and as an acid tartrate of potass in tamarinds, pine-apples, ripe grapes, and other fruits. When pure, it is colorless, inodorous, and very sour to the taste.

12. If tartaric acid be boiled with nitric acid, it absorbs oxygen, and becomes converted into oxalic acid.

If it be boiled with sulphuric acid, it loses oxygen, and is changed into acetic acid.

When tartaric acid is heated to 400° F. with excess of caustic potass, KHO, *acetate* and *oxalate* of potassium are formed, thus—

$$H_2,C_4H_4O_6 + 3KHO = K(C\,H_3O_2) + K_2(C_2O_4) + 3H_2O.$$

13. Tartaric acid is used by calico printers to evolve chlorine from solution of bleaching powder, and produce white patterns on a colored ground by *discharging* the color in certain parts.

14. The most important compound of tartaric acid is the **Acid- or Super-tartrate of Potassium,** $KH,C_4H_4O_6$, which is deposited, in an impure state, from wine or grape juice in the act of fermenting, and forms an incrustation upon the bottoms and sides of the casks in which new wine is stored. When purified by boiling in water, (with pipeclay and animal charcoal to remove the coloring matter), and subsequent crystallization, it forms a white salt known as *cream of tartar*.

It is from this salt that tartaric acid is generally obtained, by converting it into tartrate of calcium, and then removing the lime by sulphuric acid, as shewn in § 4.

15. **Tartrate of Potassium and Sodium,** $KNa(C_4H_4O_6)$, is made by neutralizing a hot solution of cream of tartar with carbonate of sodium; crystals are deposited on evaporation, having a saline taste, and used in medicine as a purgative under the name of *Rochelle salt.*

16. **Tartrate of Antimony and Potassium,** or *tartar-emetic,* $K(SbO)(C_4H_4O_6)$, is made by boiling oxide of antimony with cream of tartar. It is a white salt, slightly efflorescent, and much employed as an emetic.

Ex. b.—Half fill two test glasses with solution of tartar-emetic, add to one some dilute sulphuric or nitric acid, the potass will be separated, and a mixture of cream of tartar and oxide of antimony will be precipitated.

Add to the other glass a solution of caustic potass or ammonia, oxide of antimony will be thrown down, soluble in excess of potass.

17. **Gallic acid.** $C_7H_6O_5$, and **Tannic acid,** $C_{27}H_{22}O_{17}$,—These acids are closely connected with each other, and are largely diffused throughout the vegetable kingdom, but especially in the bark of the oak and in nut-galls: whence the name *gal*lic acid.

18. *Tannic* acid derives its name from its use in the process of *tanning*. When a piece of skin previously cleaned and soaked in lime-water is immersed in an infusion of oak-bark or nut-galls, it absorbs the whole of the tannic acid, and forms a compound insoluble in water commonly known as *leather.*

The gallic acid present in the infusion is left behind.

19. When exposed to the air, or treated with dilute sulphuric or hydrochloric acid, tannic acid becomes changed into Gallic acid and Glucose (grape-sugar). Thus—

$$\underset{\text{Tannic Acid.}}{C_{27}H_{22}O_{17}} + 4H_2O = \underset{\text{Gallic Acid.}}{3C_7H_6O_5} + \underset{\text{Glucose.}}{C_6H_{12}O_6}.$$

20. A strong infusion of nut galls is used as a test in chemical analysis, especially for the salts of iron, with which a black or bluish black precipitate is formed, the basis of common writing ink.

See *Chap.* xxix. *Ex. d.*

Cloth or calico may be dyed of various shades of black, grey, or violet, by immersing it, first, in peracetate or persulphate of iron, and afterwards in a decoction of madder and logwood. A tanno-gallate of iron is formed in this case also.

21. As Ammonium, though a compound body, was ranked among the simple metals, because it acted like a metal in the various combinations it formed with other bodies, so there are certain organic compounds which resemble in character the simple *salt-radicles,* chlorine, iodine, &c., and *combine with other elements just as if they were themselves elements.* The most important of these compound salt-radicals is *Cyanogen.*

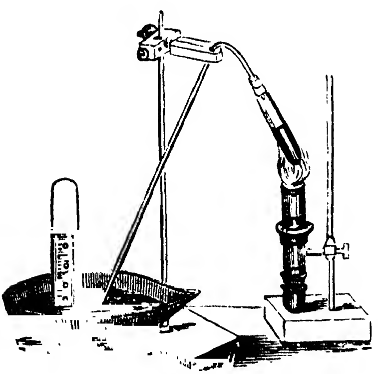


Fig. 60.

22. **Cyanogen,** **CN** =26, may be obtained by heating cyanide of mercury in a tube, when a colorless, poisonous gas will be evolved, having an odor like bitter almonds, and soluble in water, so that it must be collected over mercury. It can be condensed into a colorless, transparent liquid. Its density as compared with air is 1.806.

The symbol Cy is often used instead of CN.

Cyanogen is inflammable, and burns with a rose-colored

flame; the products of combustion are carbonic acid and nitrogen.

This may be shewn by heating a little cyanide of mercury and inflaming the gas, see fig. 81.

23. **Hydrocyanic or Prussic Acid**, HCy, is a hydracid, consisting of hydrogen and cyanogen; and may be prepared by distilling cyanide of mercury with hydrochloric acid.

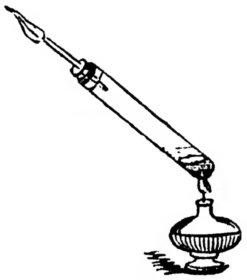


Fig. 81.

$$HgCy_2 + 2HCl = 2HCy + HgCl_2.$$

It is the most violent of known poisons. Even in *very small* quantities its vapor occasions faintness and head-ache. The most effectual antidote is ammonia, and the application of cold water, as in a shower bath.

Though used in medicine, it is in exceedingly small quantities.

"The strongest solution of prussic acid used medicinally in England, only contains 3 per cent. of real acid, and of the acid, thus diluted, a single drop is a dose."

The kernels of peaches and other stone fruits, and especially those of bitter almonds, as well as the leaves of the laurel, contain hydrocyanic acid. Hence laurel-water, and oil of bitter almonds, are poisonous.

24. When hydrocyanic acid is added to a metallic oxide, the resulting salt is a cyanide of the metal. Thus **Cyanide of Mercury**, $HgCy_2$, may be formed by dissolving *red oxide* of mercury in hydrocyanic acid, evaporating and crystallizing.

25. **Cyanide of Potassium**, KCy, or K,CN, may be obtained by heating potassium in the vapor of cyanogen, or hydrocyanic acid; in the latter case with evolution of hydrogen.

Also, by passing pure nitrogen gas through a white hot porcelain tube containing carbonate of potassium and charcoal, when carbonic oxide is given off, while cyanide of

potassium settles in the cooler portions of the tube as a white *amorphous* powder.

$$K_2CO_3+4C+2N = 3CO+2K(CN).$$

If *nitrogenous substances*, such as horn-shavings, hide-parings, hoofs of animals, &c., be heated to redness with carbonate of potassium in a close vessel, abundance of cyanide of potassium will be formed, arising from the mutual decomposition of the ammonia and the alkaline carbonate in the presence of charcoal.

$$2NH_3+K_2CO_3+C = 3H_2O+2KCN.$$

26. Cyanide of Potassium is now made in considerable quantities from the ferrocyanide, § 29, being much used in electro-plating and gilding, as a solvent for the precious metals, by forming **Cyanide of Gold**, $AuCy_3$, and **Cyanide of Silver**, $AgCy$.

Silver Solution.—Add cyanide of potassium to nitrate of silver, cyanide of silver will be deposited. When this precipitate has been well washed, a further solution of cyanide of potassium is added, which completely dissolves it; forming a double cyanide of potassium and silver.

$$AgNO_3+2KCy = KNO_3+AgKCy_2.$$

Gold Solution.—Dissolve gold in *aqua regia, Chap.* xxv. § 26, and boil the resulting perchloride of gold with calcined magnesia to precipitate the oxide. (*a*). Separate the oxide by filtration, boil it with nitric acid to remove any remaining magnesia, and after it has been well washed, dissolve it in cyanide of potassium, (*b*). when a double cyanide of potassium and gold will be formed.

(*a*.) $2AuCl_3+3MgO = 3MgCl_2+Au_2O_3.$

(*b*.) $Au_2O_3+7KCy = 3K_2O+Au_2KCy_7.$

27. Cyanogen forms three acid oxides with oxygen, which are *isomeric* bodies, that is, they contain the same proportions of the same constituents, but are yet quite dissimilar in their properties. They are

Cyanic Acid	$HCyO$	or,	$HCNO$,	*monobasic.*
Fulminic „	$H_2Cy_2O_2$	„	$H_2Cy_2N_2O_2$,	*bibasic.*
Cyanuric „	$H_3Cy_3O_3$	„	$H_3Cy_3O_3$,	*tribasic.*

The distinct character of these acids may be seen by comparing their silver-salts.

Cyanate	of Silver	AgO,CyO	$=$	$AgCyO$.
Fulminate	,, ,,	$2AgO,Cy_2O_2$	$=$	$Ag_2Cy_2O_2$.
Cyanuret	,, ,,	$3AgO,Cy_3O_3$	$=$	$Ag_3Cy_3O_3$.

28. **Fulminic acid** has never been obtained in a separate state. It derives its name from the fulminating properties of its metallic salts. **Fulminate of mercury,** $Hg_2Cy_2O_2$, is manufactured on a large scale for the purpose of filling percussion caps; it is sometimes mixed with a little sulphur and chlorate of potass, and the mixture when pressed into the cap is secured by a drop of varnish.

Fulminate of mercury is dangerously explosive, and the fulminates of silver and gold still more so. They had, therefore, better not be experimented with.

29. When a solution of cyanide of potassium is gently heated with iron filings for some time in an open vessel, oxygen is absorbed, and the iron dissolved. On evaporation, the yellow liquid obtained deposits beautiful lemon-colored crystals, tough and difficult to be powdered, and becoming anhydrous at a gentle heat. This salt is termed **Ferrocyanide of Potassium,** or *prussiate of potass.* $4K,Fe,6CN+3H_2O$, or, more simply, K_4FeCy_6.

Ferrocyanide of potassium is made on a large scale by fusing refuse animal matters with *pearlash* (impure carbonate of potassium), and scraps of old iron; cyanide of potassium is formed as stated in § 25. The fused mass is afterwards heated in water for some time, and the iron is dissolved, as shewn above.

30. The Ferrocyanides, $4M,FeCy_6$, or $4M,Cfy$, are considered to be compounds of bases with a *still more complex salt-radical* than cyanogen, to which the name Ferrocyanogen, $FeCy_6$, or Cfy, has been given, as it contains six atoms of cyanogen, and *one of iron.* It has never yet been obtained in a separate state. Ferrocyanide of potassium, K_4FeCy_6, therefore may be described as consisting of **Ferrocyanic Acid,** $H_4(FeCy_6)$ a tetrabasic acid, in which the 4 atoms of hydrogen have been replaced by 4 atoms of potassium.

31. **Ferrocyanide of Potassium forms a valuable test for metals**; when mixed with most of the metallic salts in solution, it throws down as a precipitate the ferrocyanide of each metal, and frequently with a characteristic color.

When added to a salt of iron, it forms the well-known substance termed Prussian blue; it is therefore much used by dyers, calico-printers, and color-makers.

Ex. c.—Fill a series of glasses with solutions of metallic salts, and add to each a few drops of a solution of ferrocyanide of potassium, ferrocyanides of the respective metals will be thrown down, thus—

Solutions of		*Color of precipitates.*
Iron	proto-salts.	white or pale blue.
,,	per-salts.	dark blue.
Copper	proto-salts.	lilac.
,,	per-salts.	deep brown.
Manganese	proto-chloride.	reddish white.
Lead	nitrate.	white.
Silver	,,	cream-color.
Mercury	chloride.	white.
Cobalt	,,	pale green.
Tin	,,	pale yellow.
Zinc	,,	white.
Nickel	sulphate.	gray.

Ex. d.—Heat ferrocyanide of potassium with eight or ten times its weight of strong sulphuric acid, pure carbonic oxide will be evolved. The reaction is explained, *Chap.* xi. § 21.

32. **Ferricyanide of Potassium**, sometimes called ***red prussiate of potassium***, is a sub-salt, consisting of $3K,FeCy_6$. It forms a bright blue, known as Turnbull's blue, with ***proto***salts of iron, and gives ***no*** precipitate with ***per***salts.

$$\text{Turnbull's blue} = Fe_3(FeCy_6)_2.$$

33. When anhydrous ferrocyanide of potassium, K_4FeCy_6, is fused with half its weight of sulphur, each atom of cyanogen combines with one atom of sulphur, giving rise to a new *monobasic* radical, **Sulphocyanogen**, CNS_2, or

Csy, which remains in union with the potassium and iron, forming sulphocyanides of those metals; thus—

$$K_4FeCy_6 + 6S = 4KCsy + Fe2Csy.$$

34. A soluble sulphocyanide gives no precipitate with the persalts of iron, but causes their solutions to assume a blood-red color; hence sulphocyanide of potassium is sometimes used as a test for *per*oxide of iron.

Ex. e.—Moisten a white plate with a little perchloride or persulphate of iron, and let fall on it a drop or two of sulphocyanide of potassium, it will appear exactly like blood.

EXERCISES ON CHAPTER XXXIV.

1. What are organic chemical compounds?
2. What are the prevailing elements in organic compounds?
3. What is meant by *nitrogenous* and *non-nitrogenous* compounds, and how may they be distinguished?
4. How are organic compounds divided? How may organic acids be prepared?
5. Describe the nature and properties of acetic acid.
6. Give an account of acetate of lead, and its uses.
7. How is acetate of copper prepared?
8. State the nature and properties of oxalic acid.
9. For what is oxalate of ammonia useful as a test?
10. Why is oxalic acid used to remove ink stains?
11. What is tartaric acid?
12. How is tartaric acid acted on by nitric and sulphuric acids, and by caustic potass?
13. For what is tartaric acid used?
14. What is *cream of tartar*, and for what is it used?
15. What is *Rochelle salt?*
16. What is *tartar-emetic?*
17. Where are gallic and tannic acids found?
18. Explain the principle of the process of tanning skins, &c.
19. How is tannic acid altered by exposure to the air.
20. What is common writing-ink? How are ordinary black dyes produced?
21. Why may cyanogen be classed with the elements?
22. Describe the nature and properties of cyanogen.
23. How may hydrocyanic acid be obtained, and what is its most characteristic property?
24. Why should the compounds of hydrocyanic acid be classed among the haloid salts?
25. How may cyanide of potassium be procured?

26. For what is cyanide of potassium principally used?
27. Give an account of the acid oxides of cyanogen.
28. For what is fulminating mercury used?
29. What is ferrocyanide of potassium, and how is it prepared?
30. What is the constitution of the ferrocyanides?
31. For what is ferrocyanide of potassium used?
32. What are the properties of ferridcyanide of potassium?
33. How is sulphocyanogen obtained?
34. For what is sulphocyanide of potassium a test?

CHAPTER XXXV.

ORGANIC COMPOUNDS—*continued.*

1. Some of the most important organic *bases* in the vegetable kingdom are used in medicine, and from their power of combining with acids have been termed *alkaloids* or *vegeto-alkalies.* They are, **Quinine**, the active principle in Peruvian bark; **Morphine**, the highly poisonous principle in opium and laudanum; **Strychnine**, a powerful poison usually prepared from *nux-vomica.* There are also **Theine**, from tea; **Nicotine**, from tobacco; and many others.

2. Many of the neutral organic bodies are of great importance in the arts and medicine, and compose the principal part of the food of man and animals. They are, as already stated, *Chap.* xxxiv. § 3, divisible into *nitrogenous* and *non-nitrogenous* bodies. The former contain the most nutriment; the latter contribute little to the actual growth of animals, but are, as it were the fuel, by the combustion of which (that is its union with oxygen) the natural warmth of the body is maintained.

3. When more non-nitrogenous food is taken than is required for this purpose, it is converted into fat, which consists chiefly of carbon and hydrogen; when too little is taken, the body becomes lean, because the already existing fat is absorbed to maintain the vital warmth.

Hence, animals that sleep through the winter, though fat at the commencement of their period of hybernation, become very lean towards its close.

For a like reason, the inhabitants of cold climates require a larger proportion of fat, oily, substances in their food, than those of more temperate regions.

4. The chief *non-nitrogenous* compounds gum, sugar, woody fibre or lignine, fats, oils, and resins. They all consist of charcoal and the elements of water.

5. **Starch**, $C_6H_{10}O_5$, is found abundantly in all plants, and is obtained,—from their seeds or fruit, as wheat-starch, or rice-starch;—from their roots or underground stems, as potato-starch, tapioca, and arrow-root;—from the pith of certain trees, as sago, &c.

6 By the action of sulphuric acid, starch may be converted into a kind of gum, called dextrine,—or into sugar, *Chap.* iii. *Ex. q.* The latter change takes place during the first growth or sprouting of seeds; hence *malt*, which is barley kept moist and warm till it begins to grow, has a sweet taste which the barley itself has not.

7. **Sugar** is the name given to any sweet substance contained in plants or animals. Hence there are different varieties of sugar, the chief of which are:

Cane or Ordinary Sugar, $C_{12}H_{22}O_{11}$, produced from the sugar-cane, beet-root, sugar-maple, &c.

Grape Sugar, or the sugar of fruits, $C_6H_{12}O_6$, prepared also from honey and starch.

Milk Sugar, or sugar of milk, $C_6H_{12}O_6$, obtained by evaporating and crystallizing *whey*.

Glycerine, or sugar of oil, $C_3H_8O_3$, which differs from the preceding in not yielding *alcohol* by fermentation.

8. **Lignine** or woody fibre forms the structure of plants; fine linen or cotton exhibits it in a nearly pure state.

If linen rags, tow, sawdust, be exposed to the action of

sulphuric acid, as was described under the account of starch, § 6, they will be converted into gum or sugar, according to the nature of the process.

9. When lignine, in the form of cotton, muslin, sawdust, &c., is soaked for about five minutes in a mixture of equal parts of strong nitric and sulphuric acid, and then taken out, well washed and cautiously dried, it will be found unaltered in appearance, but much changed in properties. It is now **Gun-cotton**, or *pyroxyline*, $C_6H_8(NO_2)_2O_5$.

Gun-cotton is highly explosive, detonating sharply when struck with a hammer on an anvil. It burns very rapidly, without smoke or residue.

Ex. a.—Put some loose gun-cotton on the hand, or on a little gunpowder, and inflame it by touching with a heated wire. It will neither burn the hand nor ignite the powder, owing to the rapidity of the combustion.

10. Gun-cotton dissolves in ether, forming a syrupy liquid, which on evaporating leaves behind a transparent, tenacious film. This substance, termed **Collodion**, is largely employed in photography.

If collodion be spread over a wound the sides of which are held together, it evaporates, leaving a tenacious film, which serves the purposes of plaster, and is insoluble in water.

11. The chief *nitrogenous* substances are, albumen, fibrine, and caseine, which are found both in the vegetable and animal kingdoms.

12. **Albumen** occurs nearly pure in the white of eggs; it is also largely found in the blood, the humors of the eye, the flesh of animals, and in oily seeds. It is soluble in water, and is coagulated into a white, opaque mass by heat, by most acids, and by many metallic salts.

Hence white of egg is given in cases of poisoning with salts of copper or mercury, or with arsenious acid.

Ex. b—The presence of an alkali (*soda*) in albumen may be shewn by dropping a little white of egg into an infusion of vegetable blue, when the blue will be changed to green.

13. **Fibrine** is found soluble in the living blood, but after death or separation from the body, it spontaneously coagulates in fibres, which are insoluble in water. Fibrine occurs also in the lean flesh or muscular fibre, whence it has derived its name.

14. If a little wheat flour be tied up in a piece of linen rag, and kneaded under a stream of water, so long as the latter flows off milky, a tough, adhesive elastic substance will be found in the cloth, which is *gluten* or *vegetable fibrine.* It is tasteless, and odorless; hardens on drying, into a horn-like mass, and is identical in composition with the fibrine of the blood.

The milky color is given to the water by the separation of the starch which is diffused through it, and settles slowly to the bottom of the vessel.

15. **Caseine** is the *curd* of milk, and derives its name from the Latin, *caseus,* cheese, of which substance it is the chief constituent.

It is found in most plants, but especially in peas, beans, and other seeds of a similar kind.

It has the same composition as albumen and fibrine, but its solution differs from that of the former, in being coagulated by vinegar, and in not being coagulated by heat, and from that of the latter, in not coagulating spontaneously.

16. Nitrogenous substances, when in a moist condition, putrefy, or decompose of their own accord; if added to bodies of the non-nitrogenous class while in this state of decomposition they produce a change also in them,—this change is termed *fermentation.*

Pure sugar, dissolved in water, undergoes no change; but if a little *yeast* be added, and the mixture kept in a warm place, it will effervesce from the evolution of carbonic acid. The sugar has been converted into carbonic acid and alcohol; the latter substance may be separated from the watery part of the solution by distillation.

17. **Alcohol,** *spirits of wine,* C_4H_6O, is produced by

the breaking up of an atom of grape sugar into alcohol, and carbonic acid, and water ; thus—

$$C_6H_{12}O_6 = 2(C_2H_6O) + 2CO_2.$$

Grape sugar. Alcohol.

It is the grape sugar alone that yields alcohol, the ferment or yeast changing the cane sugar into that substance ; the same may be said of milk sugar.

18. Pure Alcohol is a colorless liquid, spec. grav. 0.795. It is very inflammable, burning without smoke. From its great solvent powers it is much used by the chemist, the varnish-maker, and in many of the arts.

19. All intoxicating liquors contain alcohol ; they are not confined to vegetable preparations, for the Tartars produce an intoxicating drink, termed *koumiss*, from mares' milk, by the fermentation of the sugar of milk contained in it.

Ex. a.—For the sake of experiment alcohol may be extracted from any fermented liquor, such as porter, ale, cyder, wine, &c. It is only sufficient to put the liquid to be operated upon in a retort, and apply heat beneath. The spirit rises in steam, and this is condensed by cold.

On a larger scale it is prepared by means of a still, as in fig. 82, on the opposite page, for a description of which see *Chap.* vi. § 16.

20. The fermentation which produces alcohol is termed *vinous* (wine-making) fermentation,—if allowed to go on, however, it passes into the *acetous* (vinegar-making) fermentation.

The *al*cohol, by exposure to the air, is *dehy*drogenated—that is, deprived of hydrogen, becoming converted into water, and an intermediate substance termed **Aldehyde,** (*a*). This aldehyde rapidly absorbs oxygen from the air, and is changed into acetic acid or vinegar (*b*).

$$(a.)\ C_2H_6O + O = C_2H_4O + H_2O.$$

Alcohol. Aldehyde.

$$(b.)\ C_2H_4O + O = C_2H_4O_2.$$

Aldehyde Acetic acid.

21. When alcohol is distilled with sulphuric acid a vola-

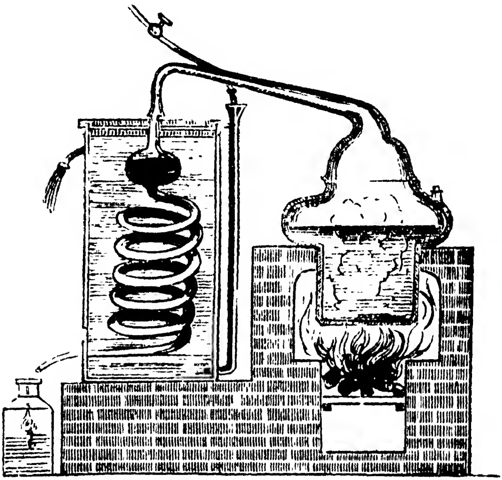


FIG. 82.

tile, fragrant liquid passes over known as **Ether**, or *sulphuric ether*, because obtained by the action of sulphuric acid.

Sulphuric Ether, or Ethyl Ether, $\left.\begin{matrix} C_2H_5 \\ C_2H_5 \end{matrix}\right\} O$, is not capable of dissolving so many substances as alcohol; still, however, it is often found useful in separating or extracting principles that are insoluble in alcohol or water, more especially in vegetable chemistry. It combines with ammonia, camphor, resins, volatile oils, sulphur, phosphorus, and chloride of gold, but has little or no action on the fixed alkalies, earths, common metallic oxides, and the greater number of the salts.

22. Various compound ethers — such as hyponitrous ether, or *sweet spirit of nitre*, &c., may be formed by the action of acids on alcohol, or by distilling a mixture of alcohol, sulphuric acid, and a salt containing the acid of the required ether.

23. When hypochlorite of lime (*bleaching powder*), alcohol, and water, are distilled together, a product comes over with the first portions of water, termed **Chloroform**, $CH\,Cl_3$. It is a dense, limpid fluid, half as heavy again as water, and very volatile. Its vapor, when breathed, produces temporary insensibility to pain; it is, therefore, often administered previous to the performance of surgical operations.

24. It would not be advisable to enter more upon the chemistry of organic bodies; the reactions that take place in the formation of ethers, &c. are somewhat intricate, and hardly fall within the range of an *elementary* work. We will, therefore, conclude what we have to say on the matter, by quoting from a French chemist the following contrast of the actions of animal and vegtable life:—

The Vegetable	The Animal
Is an apparatus of de-oxidation,	Is an apparatus of oxidation,
DISENGAGES oxygen,	ABSORBS oxygen,
PRODUCES the neutral nitrogenous bodies, sugar, starch, gum, fatty substances, &c.,	CONSUMES the neutral nitrogenous bodies, sugar, starch, gum, fatty substances, &c.,
DECOMPOSES carbonic acid, water, and ammoniacal salts.	PRODUCES carbonic acid, water, and ammoniacal salts.

EXERCISES ON CHAPTER XXXV.

1. What are alkaloids? Name the most important of them.
2. How may organic bodies be classed? What are their characteristics as articles of food?
3. What is the effect of an excess or deficiency of non-nitrogenous bodies in food?
4. Give the names and composition of the chief non-nitrogenous bodies.
5. Whence is starch obtained?
6. Why has malt a sweet taste?

7 Give the names, composition, and origin of the principal varieties of sugar.
8. What is lignine? How may sugar be obtained from rags, sawdust, &c.?
9. What is gun-cotton?
10. What is collodion?
11. Name the chief neutral nitrogenous bodies.
12. State the chief properties of albumen.
13. Whence is animal fibrine obtained?
14. How may vegetable fibrine be prepared?
15. What is caseine?
16. What is meant by *fermentation?* Give an illustration of it.
17. How is alcohol obtained?
18. What are the properties and uses of alcohol?
19. What preparations contain alcohol?
20. How does wine become vinegar?
21. How is ether prepared, and what are its properties?
22. How may the compound ethers be procured?
23. How is chloroform obtained? What are its properties?
24. Illustrate the opposing actions of animal and vegetable life.

ADDENDUM.

THE sesquicarbonate of ammonium, as given on page 210, is somewhat anomalous in its composition, with an additional atom of water it would be as regular as the corresponding sesquicarbonates of potassium and sodium; thus:—

Sesquicarbonate of	Potassium	$K_4H_23CO_3$.
,, ,,	Sodium	$Na_4H_23CO_3$.
,, ,,	Ammonium	$(NH_4)_4H_23CO_4$.

APPENDIX.

SOME chemists have adopted a different nomenclature for compound bodies from that used in the preceding pages, placing the metallic, or *quasi*-metallic, part of the compound first; thus :—

Nitric acid becomes **Hydric Nitrate.**

Potass or Oxide of Potassium becomes **Potassic Oxide.**

Lime or Hydrate of Calcium becomes **Calcic Hydrate.**

Proto-Sulphate of Iron becomes **Ferrous Sulphate.**

Nitrate of Copper becomes **Copper or Cupric Nitrate.**

Acetate of Lead becomes **Plumbic Acetate.**

Calomel, or Subchloride of Mercury becomes **Mercurous Chloride.**

Corrosive sublimate or Chloride of Mercury becomes **Mercuric Chloride.**

&c., &c.

The following tables contain the names and symbols of the principal Chemical compounds.

I. NAMES AND SYMBOLS OF THE PRINCIPAL COMPOUNDS OF THE METALLOIDS.

Water ...	... H_2O	Nitric Oxide ...	... NO
Nitrous Acid	... HNO_2	Nitrous Anhydride	... N_2O_3
Nitric Acid ...	... HNO_3	Hyponitric Anhydride ...	NO_2
Sulphurous Acid	... H_2SO_3	Nitric Anhydride	... N_2O_5
Sulphuric Acid	... H_2SO_4	Carbonic Oxide	... CO
Hydrochloric Acid ...	HCl	Carbonic Anhydride	... CO_2
Carbonic Acid (aqueous)	H_2CO_3	Sulphuric Anhydride	... SO_3
Phosphoric Acid	... H_3PO_4	Marsh Gas ...	... CH_4
Boracic Acid	... H_3BO_3	Olefiant Gas ...	... C_2H_4
Silicic Acid ...	... H_4SiO_4	Ammonia (gaseous)	... NH_3
Chromic Acid	... H_2CrO_4	Phosphoric Anhydride...	P_2O_5
Arsenious Acid	... H_3AsO_3	Sulphuretted Hydrogen.	H_2S
Arsenic Acid	... H_3,AsO_4	Phosphuretted Hydrogen	H_3P
Oxalic Acid ...	... H_2,C_2O_4	Alcohol	$\left.\begin{matrix} C_2H_5 \\ H \end{matrix}\right\} O$
Acetic Acid ...	... $H,C_2H_3O_2$		
Tartaric Acid	... $H_2,C_4H_4O_6$	Ether	$\left.\begin{matrix} C_2H_5 \\ C_2H_5 \end{matrix}\right\} O$
Nitrous Oxide	... N_2O		

II. Names, Synonymes and Symbols of some of the commoner Salts.

Caustic Potash	... Hydrate of Potassium	... KHO
Salt-petre ...	... Nitrate of Potassium	... KNO_3
Chalk ...	... Carbonate of Calcium	... $CaCO_3$
Quick-lime ...	... Oxide of Calcium ...	... CaO
Slaked-lime ...	... Hydrate of Calcium	... CaH_2O_2
Gypsum ...	... Sulphate of Calcium	... $CaSO_4$
Ammonia (liquid)	... Hydrate of Ammonium	... $(NH_4)HO$
Sal-ammoniac	... Chloride of Ammonium	... $(NH_4)Cl$
Epsom salts ...	... Sulphate of Magnesium	... $MgSO_4$
Alum ...	... Sulphate of Aluminum and Potassium ...	... $AlK2SO_4$
Sea-salt ...	... Chloride of Sodium	... NaCl
Green Vitriol	... Sulphate of Iron ...	... $FeSO_4$
Blue Vitriol	... Sulphate of Copper	... $CuSO_4$
White Vitriol	... Sulphate of Zinc ...	... $ZnSO_4$
Sugar of Lead	... Acetate of Lead ...	... $Pb(C_2H_3O_2)_2$
White Lead ...	... Carbonate of Lead ...	... $PbCO_3$
Lunar Caustic	... Nitrate of Silver ...	... $AgNO_3$
Corrosive Sublimate	Chloride of Mercury	... $HgCl_2$
Calomel ...	... Sub-chloride of Mercury	... HgCl
Tartar-emetic	... Tartrate of Antimony and Potassium ...	... $K,SbO,(C_4H_4C$
Cream of Tartar	... Acid-tartrate of Potassium .	$KH(C_4H_4O_6)$
Prussian Blue	... Ferro-cyanide of Iron	... $Fe_4,3(FeCy_6)$
Emerald Green	... Arsenite of Copper	... $CuHAsO_3$
Vermilion ...	... Sulphuret of Mercury	... HgS

III. Formulæ of Salts of Elements of different Atomicities.

	Monatomic.	Diatomic.	Triatomic.	Tetratomic
Chlorides ...	KCl	$CaCl_2$	$SbCl_3$	$SnCl_4$
Oxides ...	K_2O	CaO	Sb_2O_3	SnO_2
Hydrates ...	KHO	CaH_2O_2	$AlHO_2$	SnH_2O_3
Nitrates ...	KNO_3	$Cu2NO_3$	$Bi3NO_3$	$M4NO_3$
Sulphates ...	K_2SO_4	$CuSO_4$	Al_23SO_4	$M2SO_4$
Phosphates ...	Na_3PO_4	Ca_32PO_4	MPO_4	M_34PO_4
Carbonates ...	K_2CO_3	$CaCO_3$	M_23CO_3	$M2CO_3$

The formulæ of the Iodides, Bromides, and Fluorides are analogues to the Chlorides.

The formulæ of the Nitrates, Sulphates, and Phosphates illustrat respectively the salts formed by monobasic, dibasic and tribasic acid

IV. Examples of Chemical Calculations.

How much Chlorate of Potassium will be required to make 2 gallons of Oxygen at the ordinary temperature and pressure?

If $122\frac{1}{2}$ grains of Chlorate yield 48 grains of Oxygen.
16 grains of Oxygen = $46\frac{2}{3}$ cubic inches.
$277\frac{1}{4}$ cubic inches = 1 gallon.
2 gallons from x grains of Chlorate?

$$x = \frac{\overset{\overset{7}{\cancel{49}}}{\cancel{245}}}{\cancel{2}} \times \frac{\cancel{16}}{1} \times \frac{1109}{4} \times \frac{\cancel{2}}{1} \times \frac{1}{\underset{\cancel{8}}{\cancel{48}}} \times \frac{3}{\underset{\underset{4}{\cancel{28}}}{\cancel{140}}} = \frac{7763}{16} = 485 \text{ grains.}$$

What weight and volume of Hydrogen can be obtained by the action of dilute Sulphuric Acid on an ounce of Zinc?

If 65.2 grs. zinc yield 2 grs. Hydrogen
x grs. Hydrogen from 437.5 grs. Zinc?

$$x = \frac{437.5 \times 2}{65.2} = \frac{875}{65.2} = 13.42 \text{ grains } (\textit{weight}).$$

Again, $13.42 \times 46.6 = 625.372$ cub. in. $= 2.25$ galls. (***volume***).

What weight of marble when acted on by Hydrochloric acid yield a cubic foot of Carbonic acid gas?

100 grs. Marble yield 44 grs. Carb. Acid gas.
44 grs. C. A. gas = 93.3 cub. inches.
1728 cub. inches = x grs. of marble?

$$x = \frac{100 \times \cancel{44} \times \overset{576}{\cancel{1728}} \times 10}{\cancel{44} \times \underset{311}{\cancel{933}}} = \frac{576000}{311} = 1852 \text{ grains.}$$

V. Table of Specific Gravity of Gases and Vapou

Air = 1.000.

Hydrogen	0.069	Hydrochloric acid Gas ...	
Marsh Gas	0.557	Nitrous Oxide	
Ammonia	0.59	Carbonic Acid Gas ...	
Steam	0.62	Sulphur (vapour of) ...	
Carbonic Oxide... ...	0.967	Sulphurous Acid Gas ...	
Nitrogen...	0.972	Chlorine...	
Olefiant Gas	0.978	Phosphorus (vapour of)	
Nitric Oxide	1.04	Bromine...	
Oxygen	1.105	Mercury...	
Phosphuretted Hydrogen	1.185	Iodine	
Sulphuretted Hydrogen	1.191	Arsenic	

THE END.

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